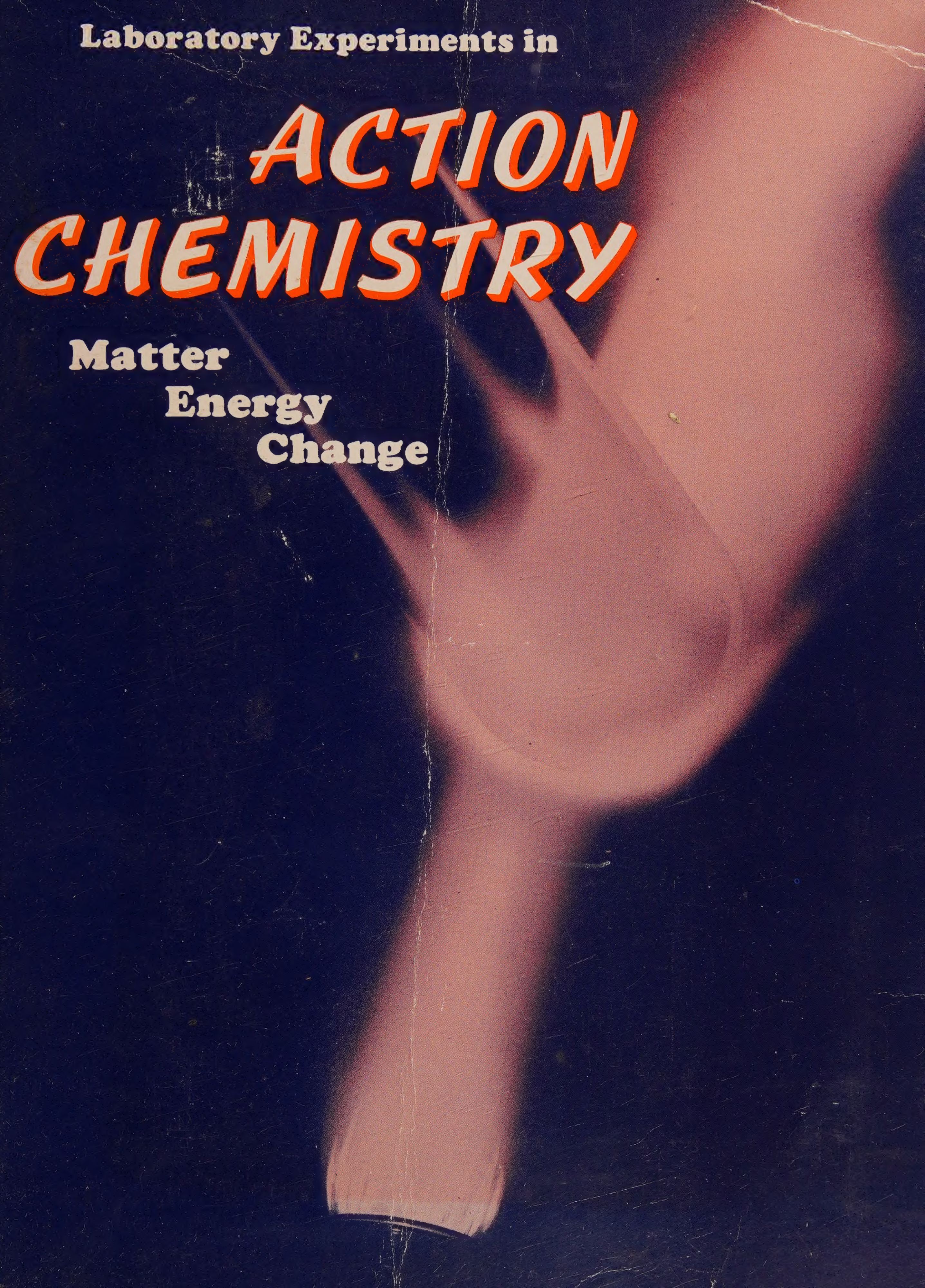
Laboratory Experiments in

ACTION CHEMISTRY

Matter
Energy
Change



Laboratory Experiments in

ACTION
CHEMISTRY

Ruth P. Bolton
Elizabeth V. Lamphere
Mario Menesini
Paul C. Huang

Holt, Rinehart and Winston, Inc.
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Preface

To the alchemist of long ago, chemistry was the stuff of magic. Today the ways of chemistry are less mysterious, but the excitement of magic is still there. In this chemistry course you will have the excitement and fun of finding out for yourself what things are made of and how they change; and that is what chemistry is all about.

The whole world is full of things to wonder about. You will cut chemistry problems down to laboratory size. By building the small ideas you discover into larger ones, you will begin to see patterns in nature. This is important because every part of your life is affected by chemistry.

You will work as scientists have always worked, asking questions, guessing at answers, trying out ideas in experiments, and wondering what the facts you discover mean. You will learn by doing and not by memorizing countless rules and definitions. What about mathematics? Of necessity there are a few number problems, but simple arithmetic will get you by. In a little while you may be surprised by the amount of real knowledge you have acquired. You may be equally amazed by the number of things you have learned to do, such as weighing and measuring and making chemistry apparatus work. Step by step you will see the many uses of chemistry in everyday life.

Ruth P. Bolton
Covina, California
July 24, 1971

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CHEMISTRY

UNITS I-IX

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Laboratory Experiments in

ACTION CHEMISTRY

UNIT I

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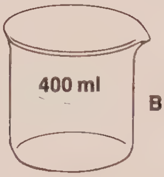
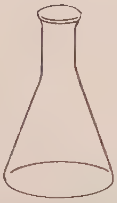
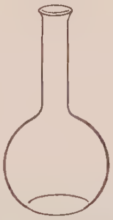
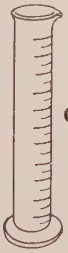
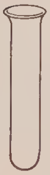
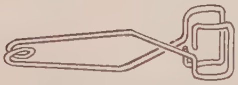
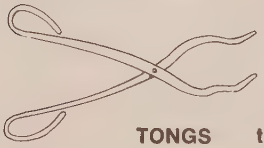
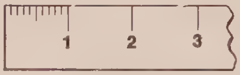
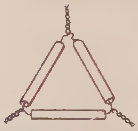
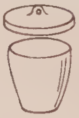
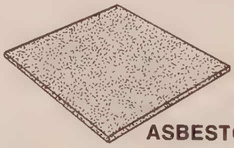
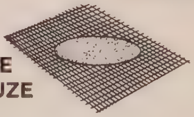
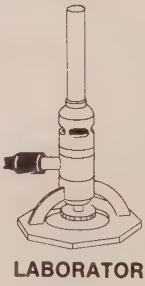
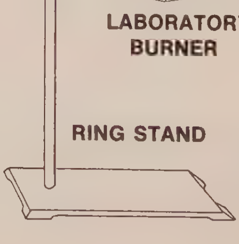
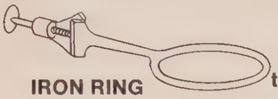
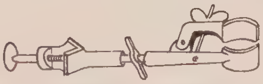
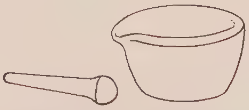
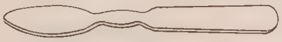
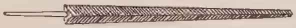
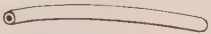
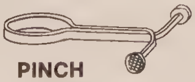
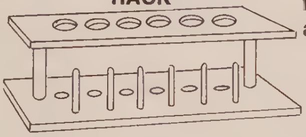
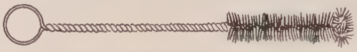
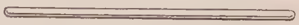
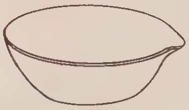
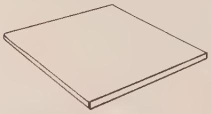
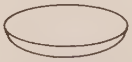
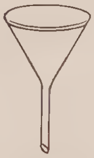
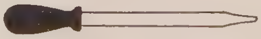
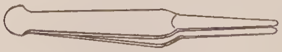
List Of Apparatus For Student Use

DESCRIPTION	APPARATUS	USE
glass common sizes 100 ml 250 ml 400 ml marked on the beaker	 **BEAKER**	as a container, like a cup may be heated
glass common sizes 125 ml 250 ml 500 ml marked on the flask	 **ERLENMEYER FLASK**  **FLORENCE FLASK**	may be heated
glass marked with a milliliter (ml) scale size — divisions 50 ml — 1.0 ml 25 ml — 0.2 or 0.5 ml 10 ml — 0.1 ml	 **GRADUATED CYLINDER**	to measure volume
glass several sizes	 **TEST TUBE**	many uses can be heated
metal clamp with a spring handle	 **TEST TUBE CLAMP**	to hold a test tube
metal	 **TONGS**	to pick up and hold apparatus

DESCRIPTION	APPARATUS	USE
10 centimeter (cm) ruler, plastic divided into centimeter and millimeter (mm) divisions	 **10 CM RULER**	to measure length
triangular wire frame with clay material coverings	 **PIPESTEM TRIANGLE**	to support the crucible
small porcelain dish with cover	 **CRUCIBLE AND COVER**	to heat small amounts of solid material at high temperature
hardened asbestos	 **ASBESTOS SQUARE**	to place under hot apparatus
wire screen asbestos center	**WIRE GAUZE** 	to spread the heat of a flame
metal heating device connected to gas outlet with rubber tubing	 **LABORATORY BURNER**	to heat chemicals in beakers or test tubes
metal rod upright heavy base	**RING STAND** 	a support with many uses

DESCRIPTION	APPARATUS	USE
iron ring with screw fastener several sizes	 IRON RING	to fasten to the ring stand as a support for apparatus
metal clamp with 1. screw fastener 2. swivel and lock nut 3. adjusting screw 4. curved clamp	 BURET CLAMP	to hold apparatus may be fastened to the ring stand
heavy porcelain dish with grinder	 MORTAR AND PESTLE	to grind chemicals to a powder
may be of metal or porcelain	 SPATULA	to transfer solid chemicals in weighing
metal file with three cutting edges	 TRIANGULAR FILE	to scratch glass to file
short length of rubber tubing	 RUBBER CONNECTOR	to connect parts of apparatus
metal clamp with finger grips	 PINCH CLAMP	to clamp a rubber connector
rack; may be wood, metal or plastic	TEST TUBE RACK 	to hold test tubes in an upright position

DESCRIPTION	APPARATUS	USE
brush with wire handle	 TEST TUBE BRUSH	to scrub glass apparatus
glass rod	 STIRRING ROD	to stir combinations of materials to use in pouring liquids
porcelain dish	 EVAPORATING DISH	as a container for small amounts of liquid being evaporated
thick glass	 GLASS PLATE	many uses (should not be heated)
curved glass	 WATCH GLASS	may be used as a beaker cover may be used in evaporating very small amounts of liquid
glass or plastic	 FUNNEL	to hold a filter paper may be used in pouring
glass tip with rubber bulb	 MEDICINE DROPPER	to transfer small amounts of liquid
metal	 FORCEPS	to pick up or hold small objects

Chapter 1

CHANGES

CHEMISTRY is the study of materials and how they change.
The laboratory is the place where we will do this study.

Experiment 1.1

Equipment and Supplies per Lab Station

beaker—250 ml or 400 ml
sandpaper; (2) test tubes; clock
(2) iron nails; marking pen or 2 labels
paper towel; (2) pieces of string; water
Solution of copper (II) sulfate

What to do

1. Take two clean iron nails. Sandpaper the nails until they are bright. Examine them. Write your observations on the chart.
2. Tie a string longer than a test tube to each of the nails.
3. You will be given two test tubes standing in a beaker. Each test tube will be one-third full of a liquid. Mark or label the tube with water: 1. Mark or label the tube with the blue copper sulfate solution: 2.
4. Lower one nail into the water. Let the string hang over the top of the test tube. Leave the nail in the water for 2 to 5 minutes. Watch carefully for any change in the nail. Using the string, remove the nail. Blot the nail gently on a piece of paper towel. Describe the results in the Data Chart. Replace the nail in the water and set the test tube in the beaker.
5. Repeat the experiment by placing the other nail in the blue copper sulfate solution in Test Tube No. 2. Let the string hang over the test tube. Leave the nail in the blue copper sulfate solution for 2 to 5 minutes. Watch carefully for any change in the nail. Using the string, remove the nail. Blot it gently on a paper towel. Describe the results in the Data Chart.

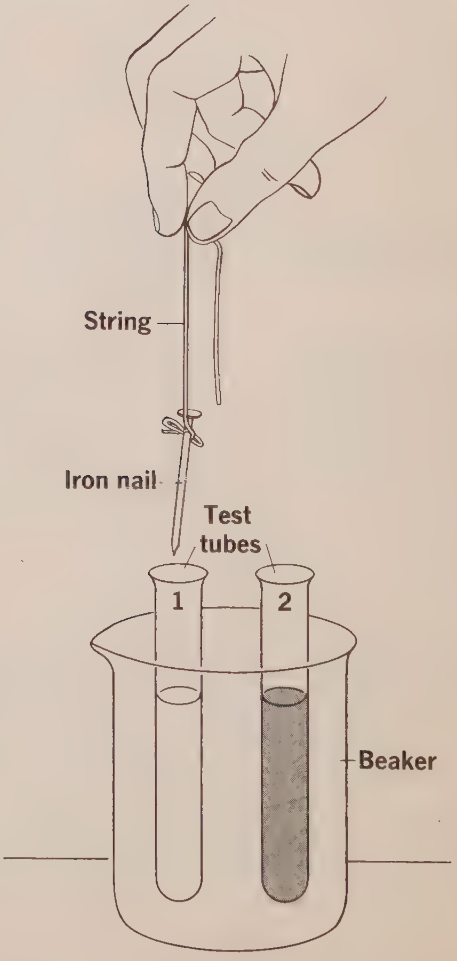
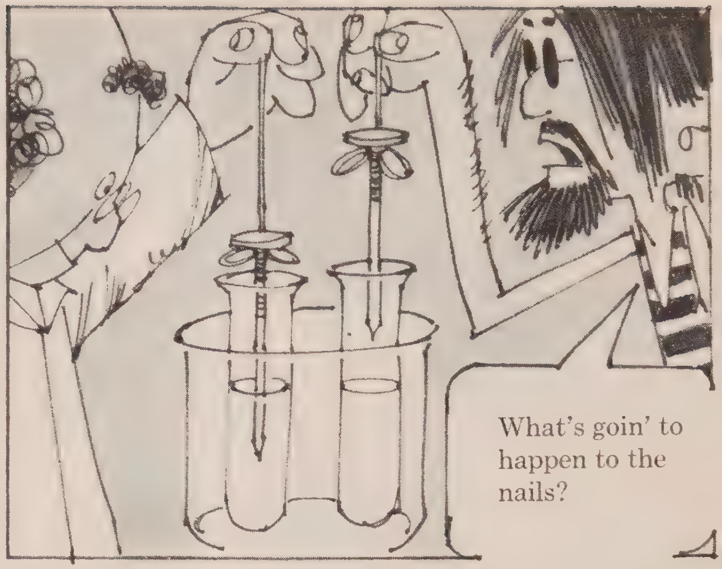


Fig. 1. Two test tubes in a beaker.

6. Return the nail to the copper sulfate solution. Let the nail slide slowly to the bottom. Place the test tube back in the beaker.
7. Write your name on a piece of paper. Place the paper in the beaker with the two test tubes. Set the beaker aside where your teacher directs since you will use the materials in Experiment 1.2.
8. Your teacher will save a sample of unused copper(II) sulfate for your use in the next experiment.


What's goin' to happen to the nails?

DATA CHART

WHAT YOU HAVE SEEN			
	Color	Dull or shiny	Other observations
Iron nail as received			
Iron nail after being suspended in water			
Iron nail as received			
Iron nail after being suspended in the blue copper(II) sulfate solution			

Questions about the experiment

1. Copper sulfate solution is made by dissolving solid copper(II) sulfate in water. Why do you think that you were told to suspend a nail in it and in the water?

   ____

   ____

2. Do you think that anything new has been formed in either test tube? ____________________

   Explain your answer. ____________________

   ____

   ____

A Little Detective Work

One of the major jobs of the chemist is to find out what is happening to substances. Let's find out what happened to your iron nail.

Experiment 1.2

Equipment and Supplies per Lab Station

beaker, test tubes, nails, and solution from Experiment 1.1
beaker—100 ml or 150 ml
test tube labeled No. 3; paper towel; sandpaper
small piece of copper foil (1 in × 1 in)
Solution of silver nitrate

What to do

1. Examine the contents of your two test tubes from Experiment 1.1.
   (a) Has there been a change in the blue liquid in Test Tube No. 2? ______________
   (b) Describe the contents of the test tubes.
   ____
   ____
2. Using its string, pull the nail out of each tube. Lay the nails on a paper towel. Describe any changes in the nails different from those you saw in Experiment 1.1.
   ____
   ____
   ____
3. Do you find any particles on the bottom of either test tube? ______________.
   If so, what do they look like?
   ____
   ____
   ____
4. Obtain a small piece of bright copper metal. How does the colored material on either nail compare with the copper metal?
   ____
   ____
   ____
5. Pour the water from Test Tube No. 1 into the sink. Now put the known copper metal into this test tube.

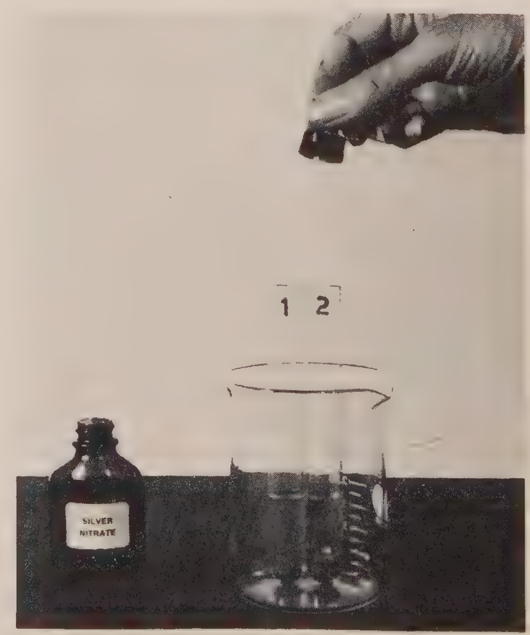
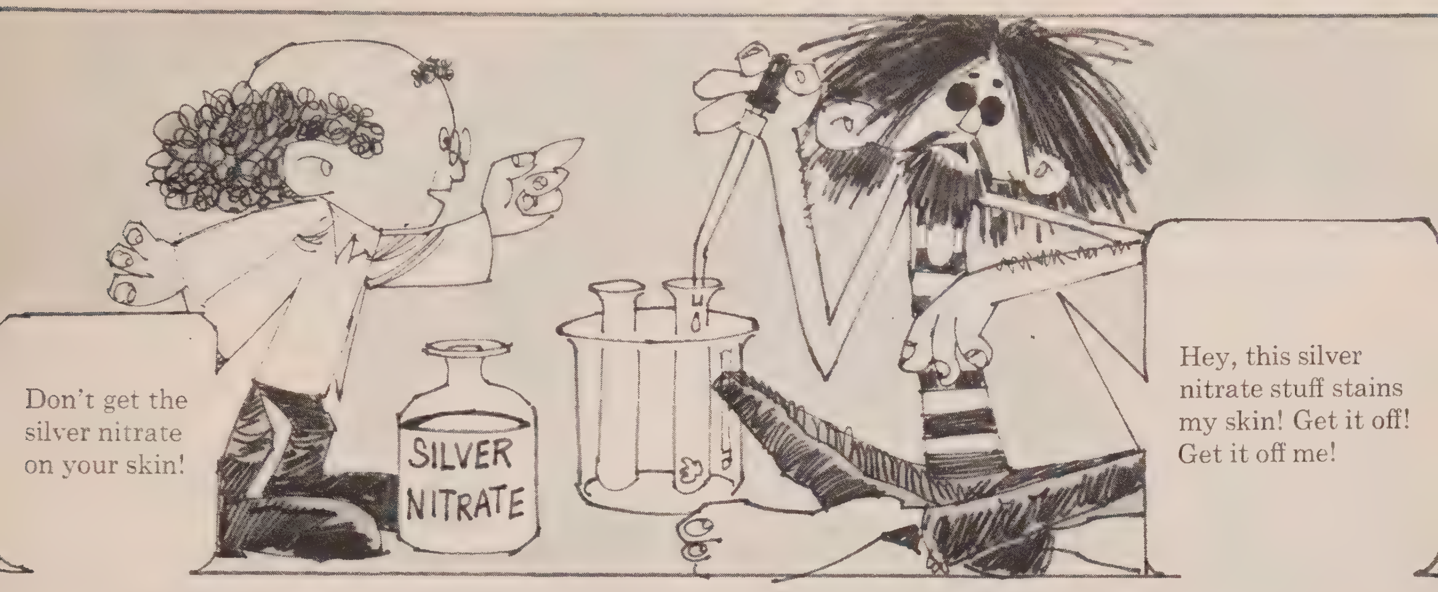


Fig. 2. Add silver nitrate.


Don't get the silver nitrate on your skin!
SILVER NITRATE
Hey, this silver nitrate stuff stains my skin! Get it off! Get it off me!

6. See How to Decant. Pour the liquid from Test Tube No. 2 into a second beaker. Be careful to leave any particles present in the bottom of the test tube. Compare the liquid in the beaker with the sample of unused copper sulfate.

______

______

7. If there is any colored material on the nail use sandpaper to scrape it into Test Tube No. 2 with possible particles already there. Place half of your particles in Test Tube No. 3. Set these particles aside for Experiment 1.2-A.

Questions about the experiment.

1. What do you believe the colored material formed on the clean iron nail to be? ______

______

______

______

8. Add a dropperful of silver nitrate solution to Test Tubes No. 1 and 2.
BE SURE NOT TO GET ANY SILVER NITRATE SOLUTION ON YOUR SKIN!

(a) What happens in Test Tube No. 1? ______

______

(b) What happens in Test Tube No. 2? ______

______

(c) Do you still agree with your answer to the second part of step #3? ______
If not, what answer would you give now?

______

______

2. What two reasons do you have for thinking this?
(a) ______

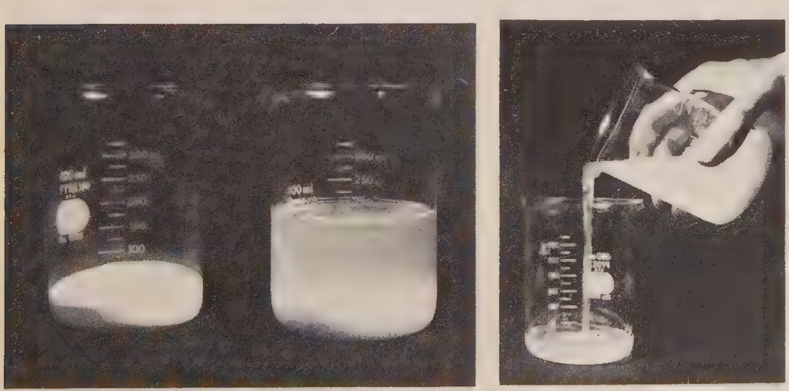
______

(b) ______

______

HOW TO

Decant a Liquid



Carefully pour off the liquid. Make sure the solid is *not* poured off.

Usetus! How do you know that the stuff left after adding silver nitrate isn't rust? Eh? Eh? How do you know? Come on!

Identity Can Be Mistaken

Mistaking a material new to you for something familiar may be easy to do.

How can you be sure that the particles found in the bottom of Test Tube No. 2 in Experiment 1.2 are not ordinary iron rust? Let's see!

Experiment 1.2-A Extension

Equipment and Supplies per Lab Station

(3) watch glasses or test tubes in a rack
white sheet of paper
sandpaper or a file, rusty nail
Solution of nitric acid, HNO_3
particles left from Experiment 1.2

What to do

1. Shake the particles you saved from Experiment 1.2 onto a watch glass.
2. Obtain a very rusty nail.
3. Scrape the rust from the nail onto a second watch glass.
4. Place a piece of copper metal on a third watch glass. Which of the three samples look most alike?

   ____

   Explain. ______________________________

   ____

   ____

   ____

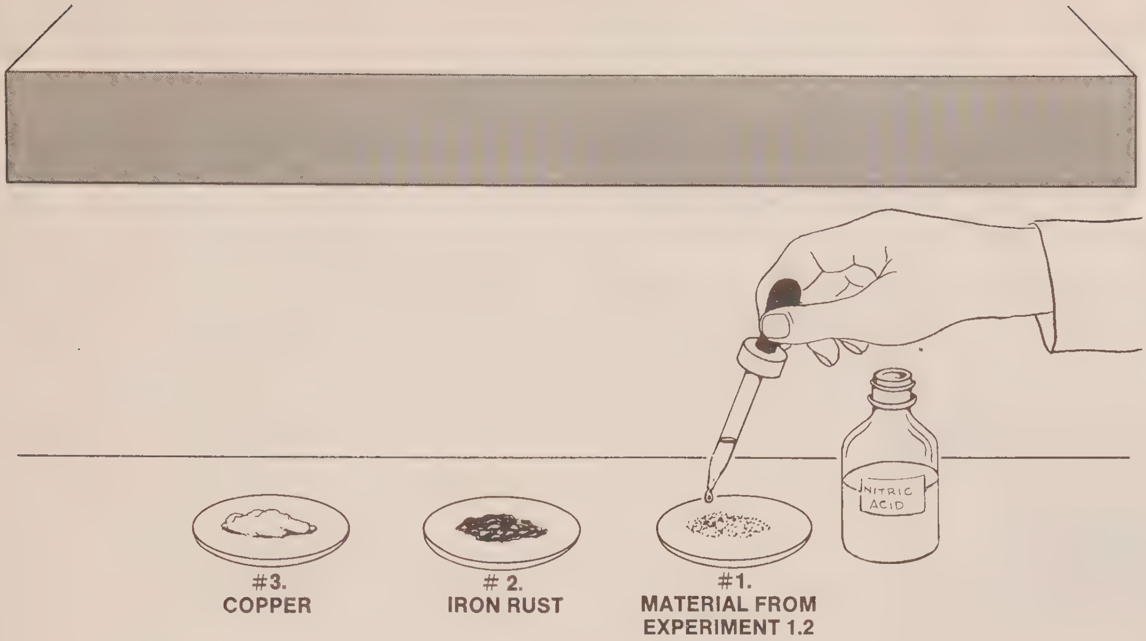


Fig. 3. Is #1 like either #2 or #3?

5. Add a dropperful of water to each sample. Place the three watch glasses on a white sheet of paper under the laboratory hood. Drop nitric acid onto each sample until each is well covered. Record your observations in the following Data Chart.

DATA CHART

NITRIC ACID TEST			
Test	Color of Solution	Material Disappears	Anything Else You Observed
1. Particles in Test Tube No. 3 from Expt. 1.2			
2. Iron Rust			
3. Copper Metal			

6. Which of the samples act in the same way with the nitric acid?

____

Questions about the experiment

1. Look carefully at your observations in this experiment.
Were the particles formed when an iron nail is left over night in copper(II) sulfate solution, copper or ordinary iron rust?
Explain. ______________________________
____

2. Experiment 1.2 and 1.2A show a method used to identify an unknown substance. First you "guess" what the substance is most likely to be, then what do you do?
____
____

TELLING IT LIKE IT IS!

In the chemistry laboratory careful records are needed. They serve as a help to the memory and as a reference for future work. Observe the changes which take place in this experiment. Record all of the changes you observe. You may use the record for this experiment in Experiment 1.4.

Experiment 1.3

PART I

Equipment and Supplies per Lab Station

beaker – dry
beaker – 250 ml half full of distilled water
stirring rod; plastic spoon
spoonful of copper (II) chloride

What to do

1. Place a level spoonful of the copper (II) chloride in a small, clean, dry beaker.

2. Fill a 250-ml beaker about half full of distilled water. Does the beaker feel warm or cool to touch?

   ____

3. Slowly add the copper (II) chloride to the water. Do NOT stir. Let stand for a few minutes. What color do you see near the crystals in the water? ____________

We'd better keep good records of what we do. Because we may need it for another experiment.

4. Stir the water with a glass stirring rod until the crystals disappear. What color is the liquid? ________ Is the beaker now cooler, warmer, or just the same as it was at first? ______________________________

PART II

Equipment and Supplies per Lab Station

test tube, forceps
aluminum foil about ½ in × 6 in

1. Get a strip of aluminum foil about ½ inch × 6 inches. Loosely wind the foil around a test tube to form a coil. Drop the aluminum into the copper(II) chloride solution.

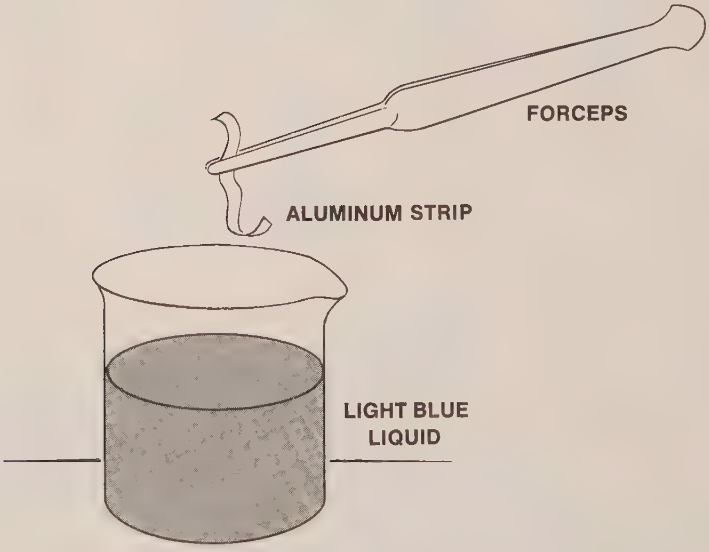


Fig. 4. Add aluminum to the blue solution.

2. Watch what happens during a period of 10 to 12 minutes. After 5 minutes cautiously touch the beaker. Is the beaker now warmer, cooler, or just the same as it was at the start in Part I?

____

3. Record the changes that you observe after the aluminum is added. There are at least eight.

____

____

____

____

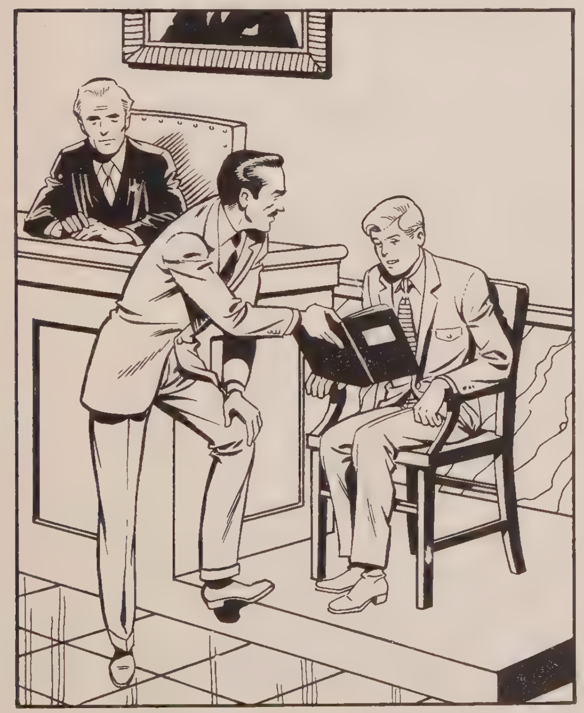
____

What happened?
Where?
Here!
Why here!
Why what?
The Aluminum?
Right!
What happened to the aluminum? And why, that's what!

LABORATORY RECORDS ARE IMPORTANT

"Record the observation. Be sure that record has the date and that you sign it."

"I see that everything is in order. Here is my stamp and signature."



"Is this the record of the laboratory upon which your company bases its claim for the patent of the new chemical substance?" "Yes, sir."

"Always remember to write down your laboratory observations and data. You never know when you may need it."

Exercise

Some of the steps which may be used in the scientific method are:

1. Identifying a problem: *Why? How? What is happening?*
2. Making a hypothesis: *A guess that must be tested!*
3. Searching the literature: *Has anyone said anything about this?*
4. Planning and doing an experiment: *We work; we test.*
5. Making observations: *We use our senses.*
6. Taking measurements: *How big? How much? How long?*
7. Recording data: *We write!*
8. Making generalizations: *There's the pattern!*
9. Making conclusions: *Here is an answer based on my observations.*
10. Communicating findings: *Let me tell you about this!*

In each of the following situations one or more of the above steps in the scientific method is being used. Which step or steps is involved? The steps do not have to be in the same order as listed above.

1. A high school student drives several second-hand cars before deciding which one he will buy.

   ____

2. A woman reads the ads in the evening paper to see which meat is on sale at the markets before she goes shopping.

   ____

3. A driver writes down the speedometer reading every time he puts gas in his car. After several weeks he figures he is getting only eight miles to the gallon of gasoline and that his carburetor needs adjusting.

   ____

4. A shopper thinks that buying the larger bottle of lotion is the best buy? After reading the bottle labels she finds that the larger bottle contains 22 oz. and costs $1.10 while the smaller bottle holds 10 oz. and costs $.48. She decides to buy the smaller bottle of lotion. Does she get the better buy? Is the smaller bottle always the better buy? ____________________
5. A rifle team member always hits the target about an inch below the bull's eye. He decides that he needs to sight a little higher before firing the gun.

   ____

WHAT'S THE ACTION?

Chemists are always asking the questions, "What?" and "Why?". The "Why?" is tougher to discover than the "What?".

In Experiment 1.3 you were discovering a few "Whats." You saw several different things happening when you put aluminum foil into a blue liquid.

In this experiment find out how aluminum acts when put into six different liquids. Three of these liquids have color. Look for likenesses and differences. These may lead you to the question "Why?".

Experiment 1.4

Equipment and Supplies per Lab Station

forceps
(6) test tubes in a test tube rack
strips of aluminum foil about 1/4 or 1/8 in. × 4 in.
(6) different liquids (of 1/2 test tubes each)

What to do

1. Number six test tubes, 1, 2, 3, 4, 5, 6.
2. Obtain the liquids listed below as your teacher may direct.

   Half fill each of the test tubes with one of the liquids as follows:
   Test Tube No. 1—distilled water
   Test Tube No. 2—yellow liquid
   Test Tube No. 3—deep blue liquid
   Test Tube No. 4—liquid containing acid
   Test Tube No. 5—green liquid
   Test Tube No. 6—liquid containing table salt

3. Drop a small piece of aluminum foil into each of the six test tubes. Watch the tests for about 10 minutes.

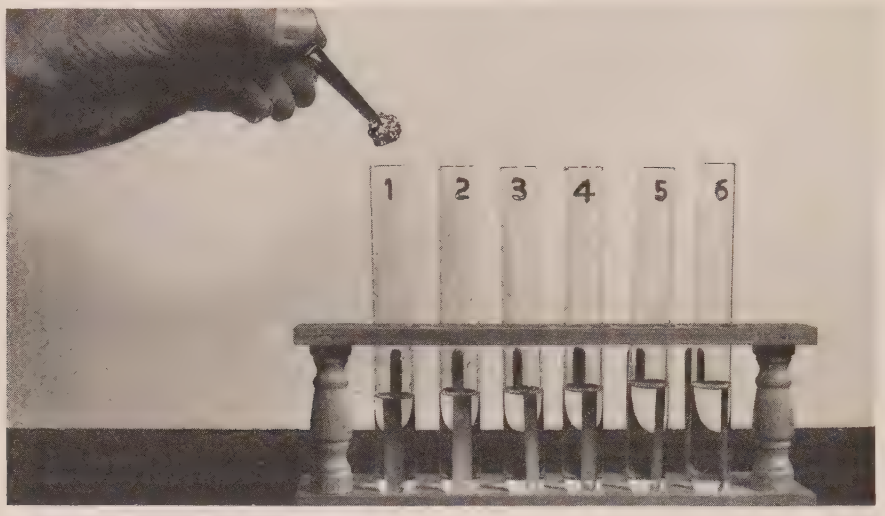


Fig. 5. Add aluminum to each of the solutions in the numbered test tubes.

4. In the table which follows, record what you see happen.

DATA CHART

ALUMINUM ADDED TO DIFFERENT LIQUIDS			
Test	Bubbles	Aluminum Disappears	Anything Else You Observed
1. Distilled Water			
2. Yellow Liquid			
3. Deep Blue Liquid			
4. Liquid with Acid			
5. Green Liquid			
6. Liquid with Table Salt			

Questions about the experiment

1. Was there bubbling in all of the colored liquids? ______
If not, which ones bubbled? ______________________
____

2. Was there bubbling in all three colorless liquids? ______
If not, which ones bubbled? ______________________
____

3. Were deposits formed in some of the tubes? ______

   ______

   If so, which ones? ______

   Was there bubbling in these tubes? ______

   If so, which ones? ______
4. Did some of the aluminum disappear in all of the liquids?

   ______ ______

   If not, in which liquids did it disappear? ______ .

   In which liquids was there bubbling? ______
5. Look over your record of what happened in this experiment and in Experiment 1.3. Do you believe that it was the colored materials in the liquids that may have caused the bubbling? ______
6. What "Why? have you asked? ______

   ______
7. Try to answer any "Whys?" you have asked.

   ______

   ______

Extra for experts

1. Vinegar is an acid. Would you store vinegar in an aluminum can? ______ Explain your answer.

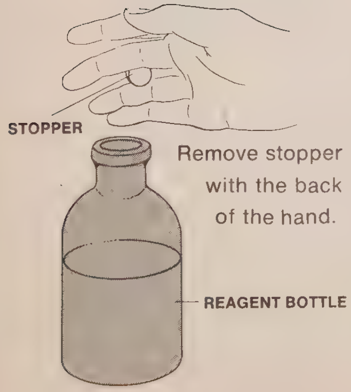
   ______
2. Why are most chemicals stored in plastic or glass containers instead of metal ones? ______

   ______
3. Aluminum cooking pots sometimes get pitted in the inside. How can you account for these tiny holes?

   ______

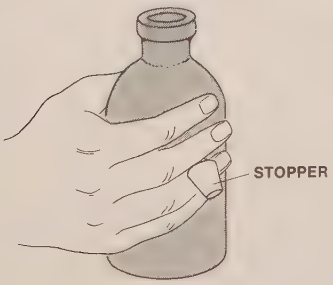
HOW TO

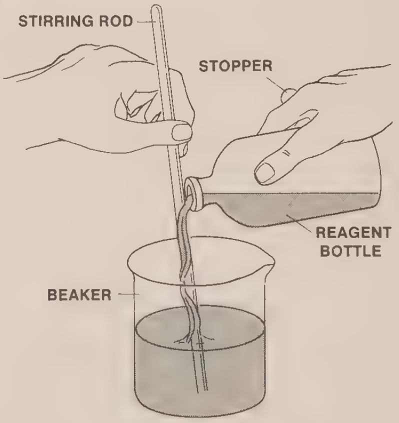
Pour Liquid From A Reagent Bottle



Remove stopper with the back of the hand.

Lift bottle making sure label is toward palm.





Pour liquid along stirring rod.

HOW TO

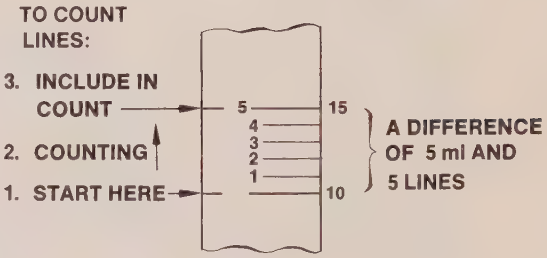
Use a graduated cylinder

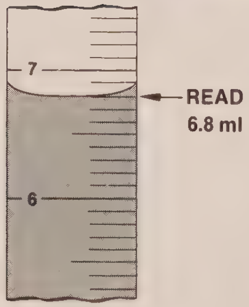
1. Decide what each unmarked line represents.

$$\text{DIVIDE:}\quad \frac{\text{difference}}{\text{number of lines}} = \frac{\text{amount}}{\text{1 line}}$$

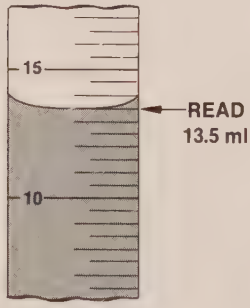
$$\text{On above scale}-\quad \frac{\text{5 ml}}{\text{5 lines}} = \frac{\text{1 ml}}{\text{1 line}}$$

Look carefully at the scale. Some graduated cylinders are marked differently.

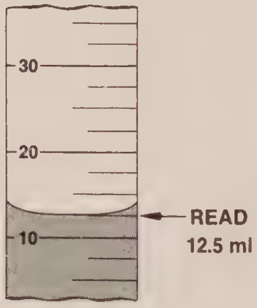




0.1 ml DIVISIONS



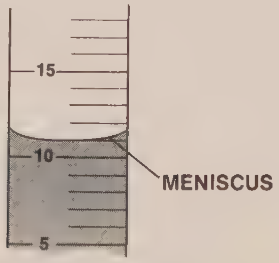
0.5 ml DIVISIONS

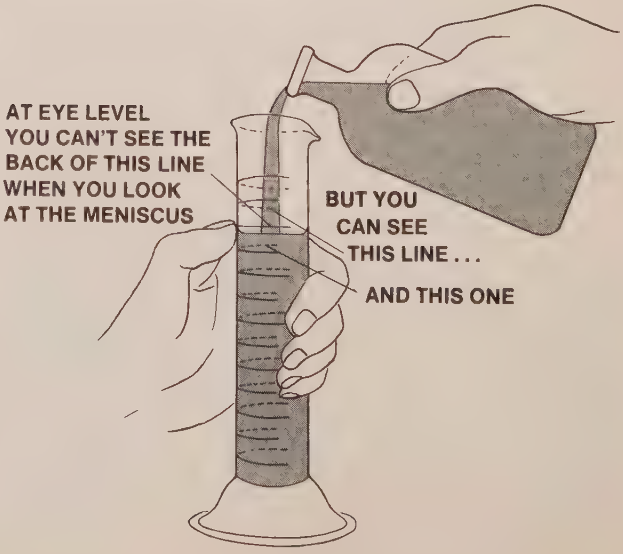


2.5 ml DIVISIONS

2. Read at the bottom of the meniscus which is the curved surface of the water.

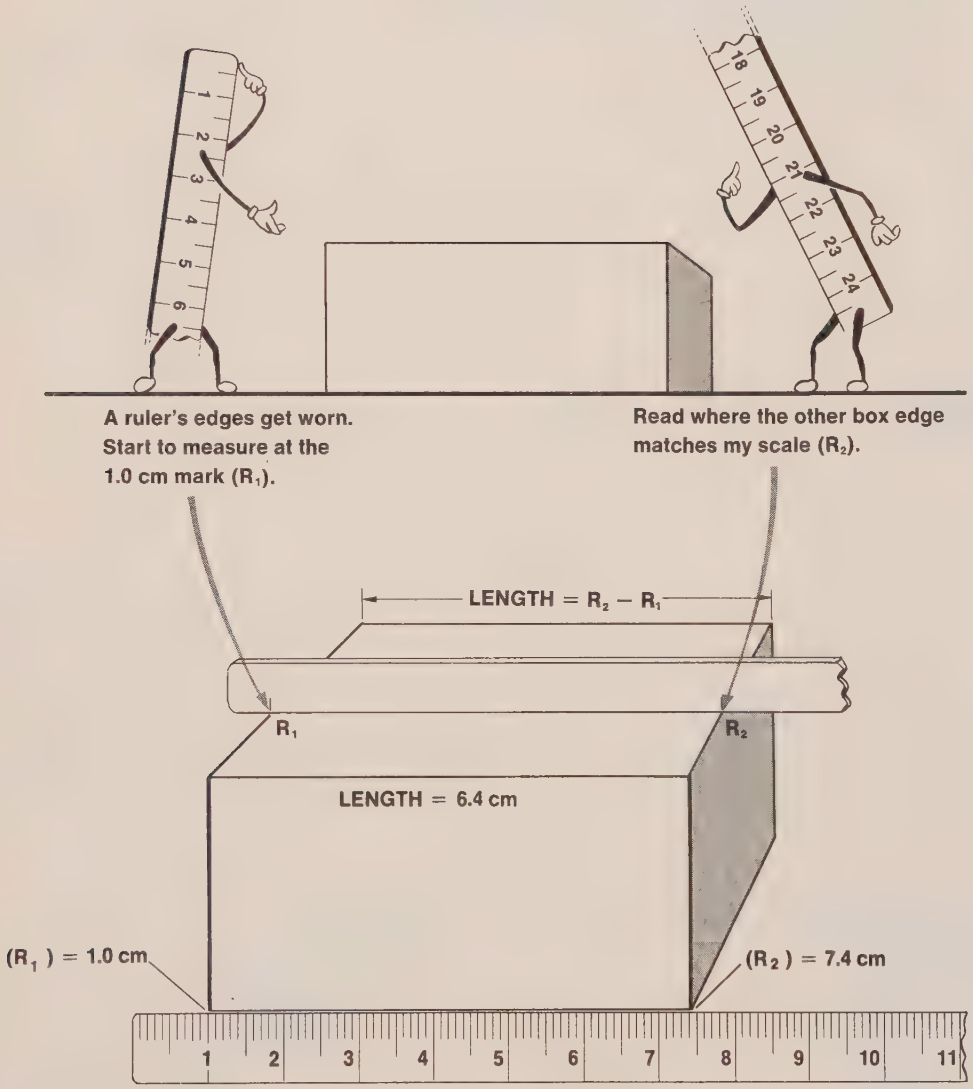
   Hold the cylinder high enough so the back of the nearest numbered line can't be seen. Then read the meniscus.




AT EYE LEVEL YOU CAN'T SEE THE BACK OF THIS LINE WHEN YOU LOOK AT THE MENISCUS
BUT YOU CAN SEE THIS LINE . . .
AND THIS ONE

HOW TO

Use a metric ruler



A ruler's edges get worn. Start to measure at the 1.0 cm mark (R_1).

Read where the other box edge matches my scale (R_2).

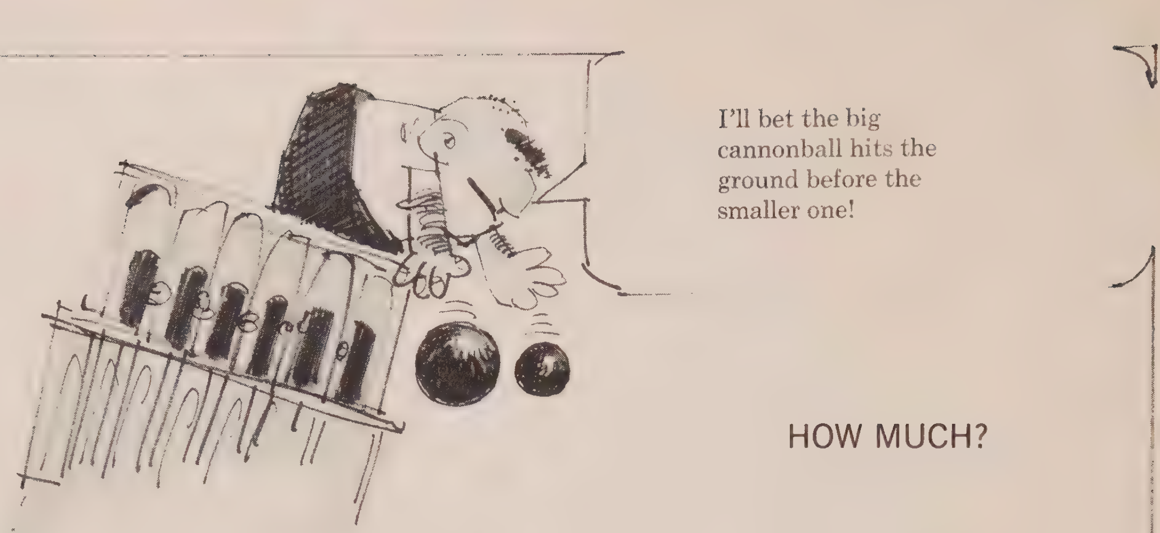
METRIC RULER

Compare this scale with the one on your ruler.

FINAL READING (R_2) MINUS INITIAL READING (R_1) = LENGTH

$R_1 = 1.0$ cm $R_2 = 7.4$ cm

7.4 cm − 1.0 cm = 6.4 cm LONG


I'll bet the big cannonball hits the ground before the smaller one!

Experiment 2.1

HOW MUCH?

How close can you guess the mass of a sample of matter using weight to help you? You can find the answer to this question with the help of your partner.

Read *page 14* of the text before starting this experiment.

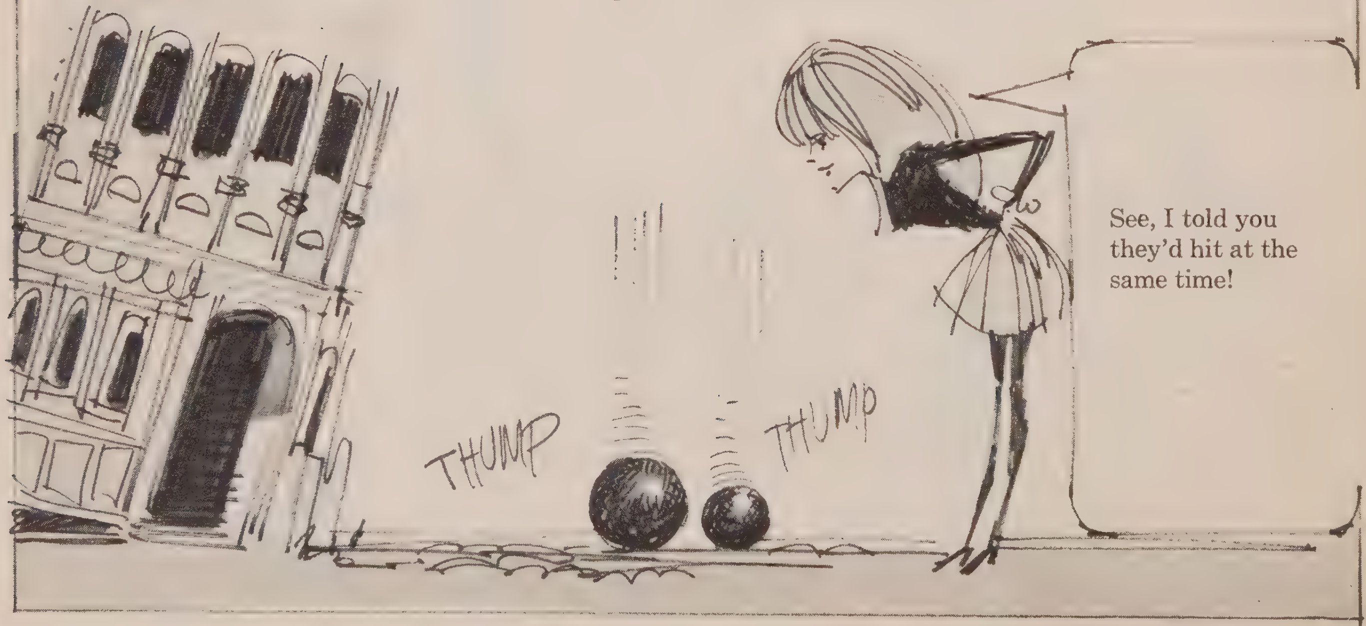
Equipment and Supplies per Lab Station

Blindfold, (2) small boxes

Set of samples (these may be shared)

What to do

1. Blindfold your partner or have him keep his eyes closed.
2. Obtain a pair of samples with the masses marked on them.
3. Have your partner hold out his hands. Place a small box in each of his hands. Then place a sample in each box. Have him compare the weights of the samples by hefting them. Remember that objects of the same mass have the same weight at a given point on earth.


See, I told you they'd hit at the same time!
THUMP
THUMP

4. Write the guesses your partner makes in the Data Chart. Then record in the Chart the masses listed on the samples.
5. Repeat steps 2–4 using other pairs of samples.

DATA CHART (My partner's observations)

Masses of Marked Samples Grams		Difference in Their Masses Grams	Partner's Identification Alike Or Not Alike
No. 1	No. 2		

6. Let your partner look at the Data Chart in your book. What is the smallest difference in mass he can detect?

   ____

7. Exchange places with your partner. Let him blindfold you, or keep your eyes closed. Your partner will repeat steps 3–5. The observations and data go in the chart in his book.
8. Look at his Data Chart. What is the smallest difference in masses that you can detect? __________

Question about the experiment

Can people always rely on their ability to measure masses accurately? ______________________________

WHAT'S ITS SIZE

VOLUME is the amount of space matter takes.
How close can you guess the volume of a sample of matter?
How can you find the volume of an odd-shaped object?
You can find the answers to these questions in this experiment.

Experiment 2.2

Equipment and Supplies per Lab Station
three different-shaped containers to hold water
graduated cylinder; paper; marking pencil

What to do

While one partner sets up Part I, the other partner does Part II. Then the partners switch. Part III may be done by both partners working together.

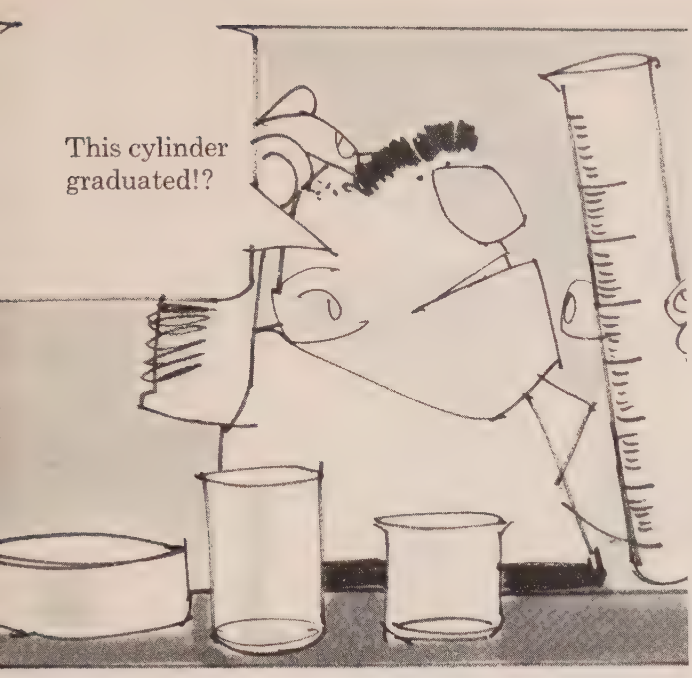
PART I

1. Obtain three or more different-shaped containers into which you can pour water.
2. Number each container.
3. The teacher will tell you what volume of water to pour into each container. Use a graduated cylinder to measure these volumes.
4. On a separate sheet of paper write down the number of the containers and the volume of water you poured into each one.
5. When your partner has finished Part II, ask him to look at the containers. Have him write in his book, the numbers of the bottles which he thinks contain the same volume of liquid.
6. Show your partner the data you wrote down in step 5, so he can complete his Data Chart and answer the question.

DATA CHART (my observations)

Numbers of containers with equal volumes of water	My guess	Partner's data

Does the shape of the container affect your ability to select equal volumes of liquid? ________________


This cylinder graduated!?

7. Go on to Part II while your partner sets up the bottles for you to inspect later. At that time you can fill in your Data Chart.

Equipment List for Part II

metric ruler
(2) boxes

PART II

What to do

1. Obtain two boxes from your teacher.
2. Inspect the two boxes. Do they appear to have equal volumes? ______________________
3. Use a metric ruler to find the height, length, and width of each box. Write the figures in the Data Chart.
4. Find the volume of each box by multiplying the length by the width by the height. The unit of volume is the cubic centimeter (c.c. or cm^3) ($cm \times cm \times cm = cm^3$).

DATA CHART

Box no.	Length (cm)	Width (cm)	Height (cm)	Volume (cm^3)

5. Do the boxes have equal volumes? ______________
How did your guess of the volumes compare with the volumes you calculated in step 4?

____

In Part III you will find another way to find the volume of a solid. Many solids are not regular in shape. We find their volume by an indirect method. In this part you will use the indirect method of finding the volumes of a neoprene stopper and a pair of forceps.

PART III

Equipment list for PART III

forceps; graduated cylinder-50 ml
neoprene stopper No. 1

What to do

1. Obtain a No. 1 neoprene stopper.
2. Half fill a 50-ml graduated cylinder with water. How many ml of water are in the cylinder? ____________
3. Carefully lower the stopper into the water in the cylinder. Gently shake the cylinder to remove any air bubbles. What is the volume of the contents of the cylinder now? ____________________

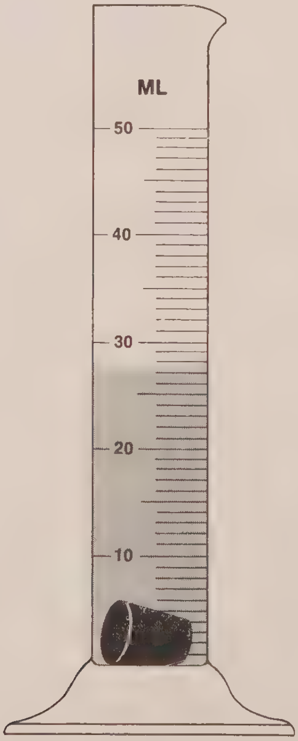


Fig. 6.

4. Calculate the volume of the stopper.

volume of water + stopper = ______________ ml

− volume of water = ______________ ml

VOLUME OF THE STOPPER = ______________ ml

5. Repeat steps 1-4 using the forceps instead of a stopper. Be sure you start with the graduated cylinder about two-third's full of water. All of the forceps must be covered by water.

6. Calculate the volume of the forceps.

volume of water + forceps = ______________ ml

− volume of water = ______________ ml

VOLUME OF THE FORCEPS = ______________ ml

Questions to answer

1. Why do you think that chemists use metric rulers, balances, and graduated cylinders to make measurements in the laboratory instead of guessing quantities? ______________________________

____

____

____

____

2. List some examples of other kinds of measurements chemists might make by using a measuring device.

____

____

____

____

____

____

____

GRADUATING A test Tube

There are other pieces of glassware besides a graduated cylinder which can be used for measuring liquids. When a piece of glassware has been marked into equal units it is said to be graduated. When the graduated apparatus is compared with a standard, then it is calibrated. We know how accurate the divisions are.

In this experiment you will make a graduated test tube and find out how accurate your markings are.

Experiment 2.2-A Extension

Equipment and Supplies per Lab Station

collection of graduated laboratory glassware
(to be shared by class)
1 each of: graduated cylinders- 10 ml, 25 ml, 100 ml
adhesive tape, pencil, test tube, and water

What to do

1. Obtain a test tube and a length of adhesive tape as long as the test tube.
2. Put the adhesive tape along the length of the test tube.
3. Measure 5 ml of water in a 25-ml graduated cylinder. Pour the water into the test tube.
4. Carefully mark on the tape a line where the bottom of the water's meniscus appears. Label the line 5 ml.
5. Repeat steps 3 and 4 by adding additional 5 ml portions of water. Mark the 10, 15, and 20 ml lines. (The capacity of your test tube will determine how much water you can add.)
6. Empty the test tube.
7. Pour water into the test tube to the 10-ml line. Then pour the water into a 100-ml graduated cylinder. How much water is in the graduated cylinder? (Read its scale for this answer.) ________

8. Measure another 10 ml of water in your test tube. Pour the water into a 100-ml graduated cylinder. How much water is in the cylinder according to its scale? ____________________

9. Measure 20 ml of water in the 25-ml graduated cylinder. Add the water to your test tube. How much water does the test tube's scale show? __________ How can you account for any differences in the measurements? ____________________
____
____

10. Save your graduated test tube for use in future experiments. You will use it to measure approximate volumes of liquid.

Questions to answer

1. Which of the four graduated pieces of glassware you used in this experiment most accurately measures a given volume of water? ____________________
Why do you think so? ____________________
____
____
____

2. Inspect other pieces of graduated glassware in the laboratory. Which of these would give the most accurate measure? ____________________
Which of these would give the least accurate measure?
____
____

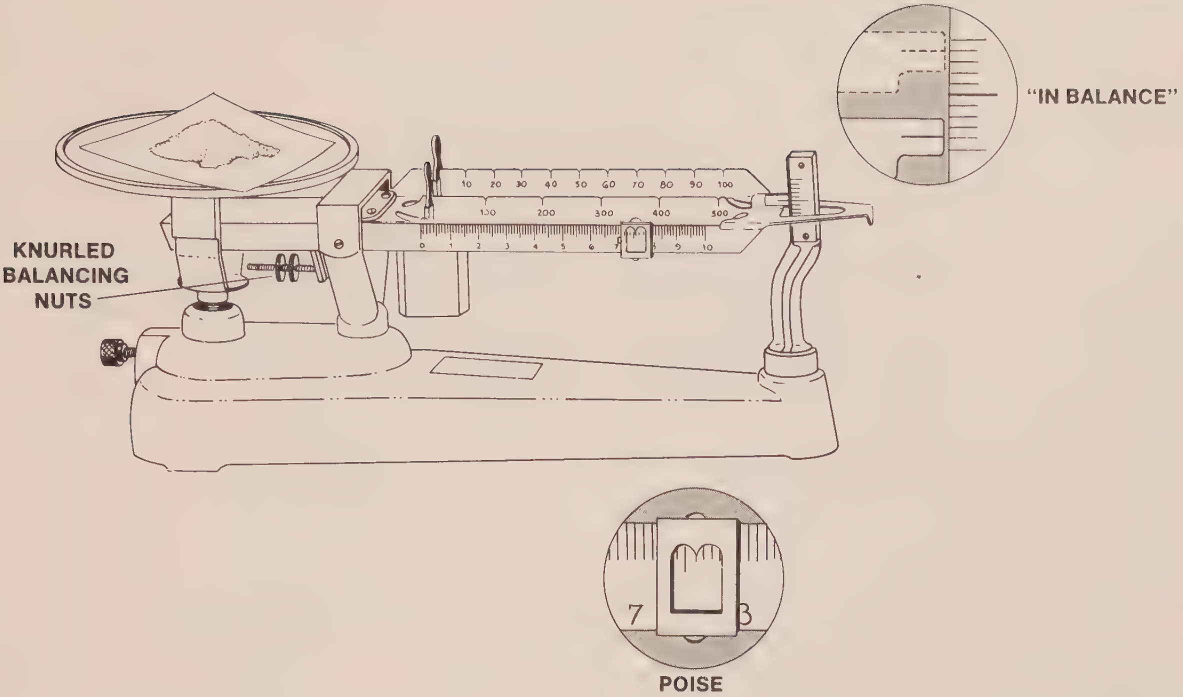
3. What basis did you use for your selection? __________
____
____
____

HOW TO USE A BALANCE

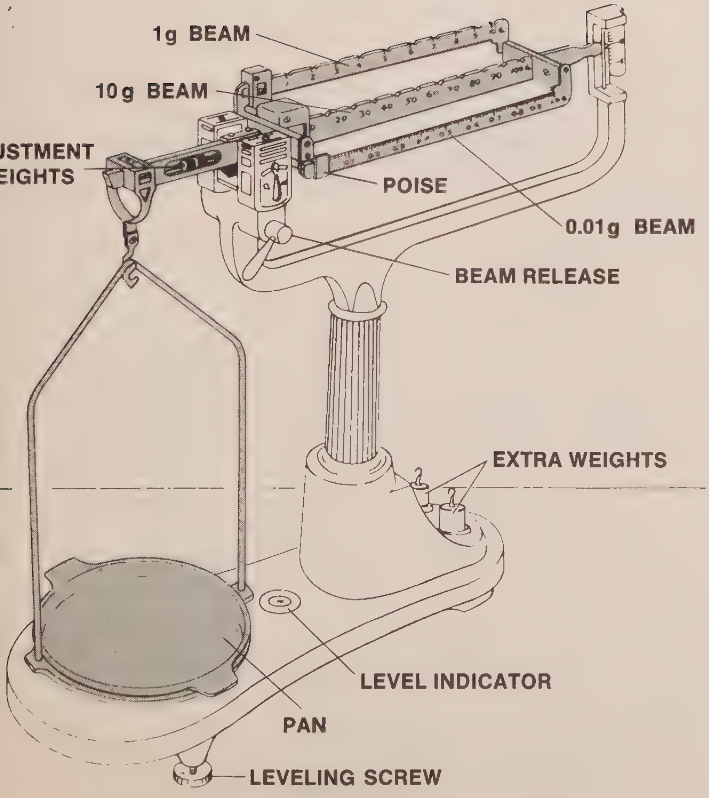
A balance is an instrument used to measure mass. There are many styles of balances. The way in which they are used depends somewhat upon the balance style. There are, however, some things which you do with most balances. In the steps below you will see how they apply to your laboratory balance.

What to do

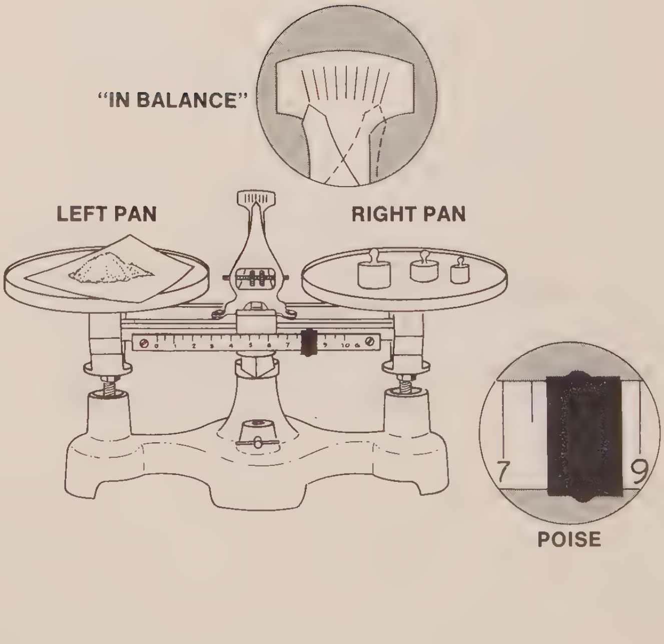
1. Look at the diagrams and find the one that most closely matches your balance. Use the diagram to help you locate on your balance any part mentioned below.
2. Set your balance on a reasonably flat and level surface.
3. Check to see if the instrument is in balance. (Look at the pointer section of the diagram to see what to expect.)
4. If the instrument is out of balance, the zero point must be adjusted. Turn the two knurled balancing nuts found on the front beam or the moveable feet at the base. Turn them until the pointer shows balance.
5. The mass to be measured is placed on the pan (the left one if there is more than one pan).
6. Move the poise along the scale until balance is restored.
7. Read the poise position on the scale. If more than one poise was used, the total of the poise scale readings is the total mass.
8. Return the poise to the zero position on the scale. Remove your mass sample.
9. Ask your teacher to help you with any of the above steps you don't know how to do.



"IN BALANCE"
KNURLED BALANCING NUTS
10 20 30 40 50 60 70 80 90 100
100 200 300 400 500
0 1 2 3 4 5 6 7 8 9 10
7
8
POISE



1g BEAM
10g BEAM
USTMENT
EIGHTS
POISE
0.01g BEAM
BEAM RELEASE
EXTRA WEIGHTS
LEVEL INDICATOR
PAN
LEVELING SCREW



"IN BALANCE"
LEFT PAN
RIGHT PAN
7
9
POISE

WHAT KIND OF MATTER IS IT?

The amount of mass which occupies 1 unit volume is called the DENSITY. You can recognize a density value by the units grams/milliliter (g/ml); grams per liter (g/l); and grams per cubic centimeter (g/cc). The density of any substance can be found by using the formula

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \qquad \text{Example: Density} = \frac{100\text{ g mass}}{50\text{ ml volume}}$$

$$\text{Density} = 100\text{ g} \div 50\text{ ml}$$

$$\text{Density} = 2.0\text{ g/ml}$$

In this experiment you will use the density value to identify an unknown substance.

Experiment 2.3

Equipment and Supplies per Lab Station

balance, metric ruler
graduated cylinder-25, 50, or 100-ml
sample-one or more

What to do

1. Inspect the substance. What do you think it is? ______________________________
2. Use the balance to determine the mass of the sample. Write the mass of the sample in the Data Chart.
3. If the sample is a rectangular solid, find its volume by doing steps 4 and 5. If the sample has an irregular shape, go on to steps 6 and 7.
4. If the sample is a rectangular solid, use a metric ruler and measure the length, width, and height. Refer to the How To Use the METRIC RULER page 19 if necessary. Write down the measurements, and then calculate the volume.

length ______ cm

width ______ cm

height ______ cm

5. Write the volume in the Data Chart.
6. If the sample is an irregular shaped object, use the displacement method studied in Experiment 2.2.
   Record the measurements:
   new height of the water level when object is immersed ______________ ml
   original height of water in cylinder without object ______________ ml
7. Subtract the two readings to calculate the volume of the sample. The volume is ______________ ml. Record this data in the Data Chart. Calculate the density of the object. Record the value in the Data Chart.

   ____

VOLUME = length × width × height

volume = ____ cm × ____ cm × ____ cm

volume of sample = ________ cm^3

DATA CHART

Sample No.	Mass	Volume	Density $= \frac{\text{Mass}}{\text{Volume}}$

8. Look at the DENSITY REFERENCE TABLE IN THE APPENDIX. What substance(s) has (have) density value(s) most like the one of your sample? ______________ What do you think your sample is? ____________________

Why? __________________________________

____

____

9. If you have time select another sample and repeat the experiment.

Sample   is identified as ______________.

Sample   is identified as ______________.

Questions to answer

1. Do you think that a big chunk of sulfur and a small piece of the same element would have the same density? ____________________

Why? ______________________________

2. What would be the density of a 350-gram sample of rock which has a volume of 100 cm^3? ______________________________

3. How would you find the density of a sample of water?

____

If your teacher approves your procedure, find the density of water. The density of water is

____

Worksheet On Density

DIRECTIONS (questions 1–3): On the line at the right, write the number of the expression which, of those given, best completes the statement or answers the question.

1. How much matter occupies a given space is known as matter's

(a) weight   (c) mass

(b) density   (d) volume   1 ______________

2. The density of water is 1 g/ml. Oil floats on water. Which density, of those given, is the oil most apt to have?

(a) 0.9 g/ml   (c) 1.9 g/ml

(b) 1.0 g/ml   (d) 10 g/ml   2 ______________

3. One hundred grams of a liquid completely fill a 200-ml bottle. The density of the liquid is

(a) 100 g/ml   (c) 0.5 g/ml

(b) 200 ml/g   (d) 2 ml/g   3 ______________

DIRECTIONS (questions 4–8) Complete the table given below:

Question No.	Mass	Volume	Density
4	40 g	5 ml	
5	1 g		1 g/ml
6		10 ml	5 g/ml
7		30 ml	3 g/ml
8	60 g		2 g/ml

9. A solution has a density of 1.5g/ml. How many grams are needed to obtain (a) 2 ml of solution (b) 5 ml of solution (c) 10 ml of solution?

(a) __________ (b) __________ (c) __________

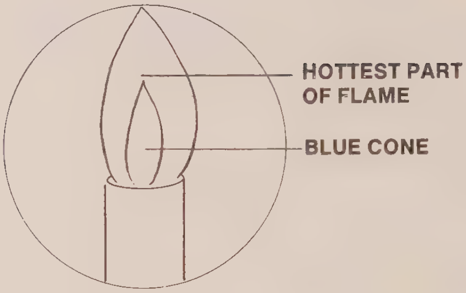
10. A solution has a density of 0.8 g/ml. How many ml are needed to obtain (a) 16 grams of solution (b) 40 grams of solution (c) 100 grams of solution?

(a) __________ (b) __________ (c) __________

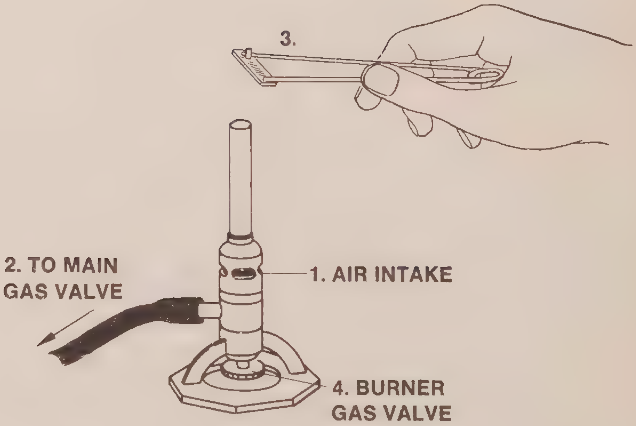
HOW TO

Light A Burner

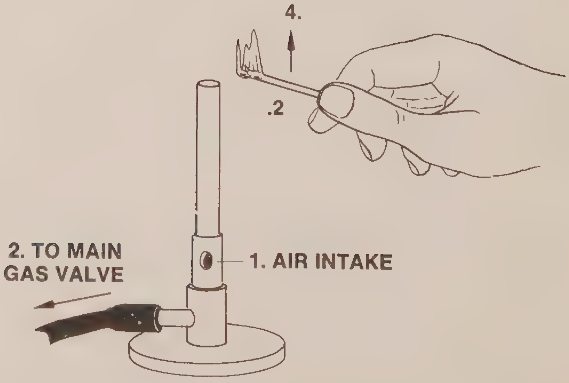
There are several styles of burners available. All have a way to adjust the air-intake. The gas intake can be adjusted at the main gas valve or by a valve at the base of the burner. The style of the burner and the source of heat to light the burner will determine the steps you follow. Listed below are procedures for the two most common burners and ways of lighting the gas.



1. Attach burner to gas.
2. Turn on gas at main valve.
3. Strike lighter and ignite gas.
4. Adjust height of flame with gas valve at burner's base.
5. Adjust the air intake to obtain a blue cone.



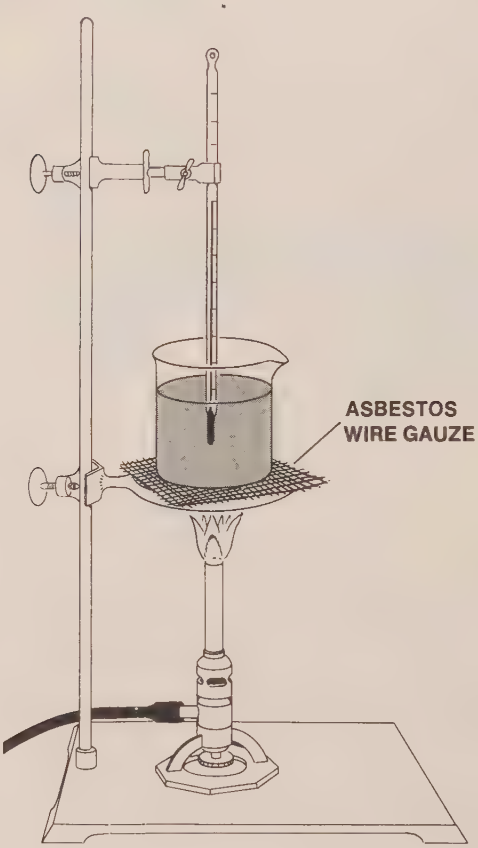
1. Attach burner to gas.
2. Strike match.
3. Turn on main gas valve.
4. Bring lighted match up the barrel to ignite gas.
5. Adjust air inlet to get blue cone.
6. Flame height must be adjusted by changing amount of gas by turning main gas valve.



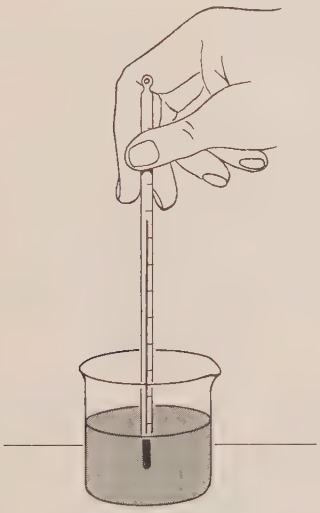
HOW TO

Use A Thermometer

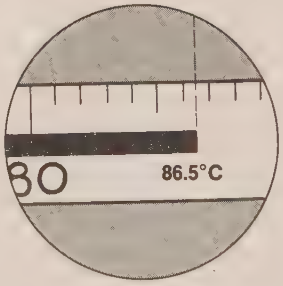
1. SELECT THERMOMETER OF RIGHT RANGE. SEE YOUR EXPERIMENT DIRECTION. BOILING POINT OF WATER = 100°C.

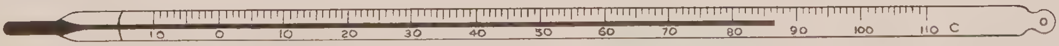


2. The bulb must not touch the container and must be in the center of the substance being heated.

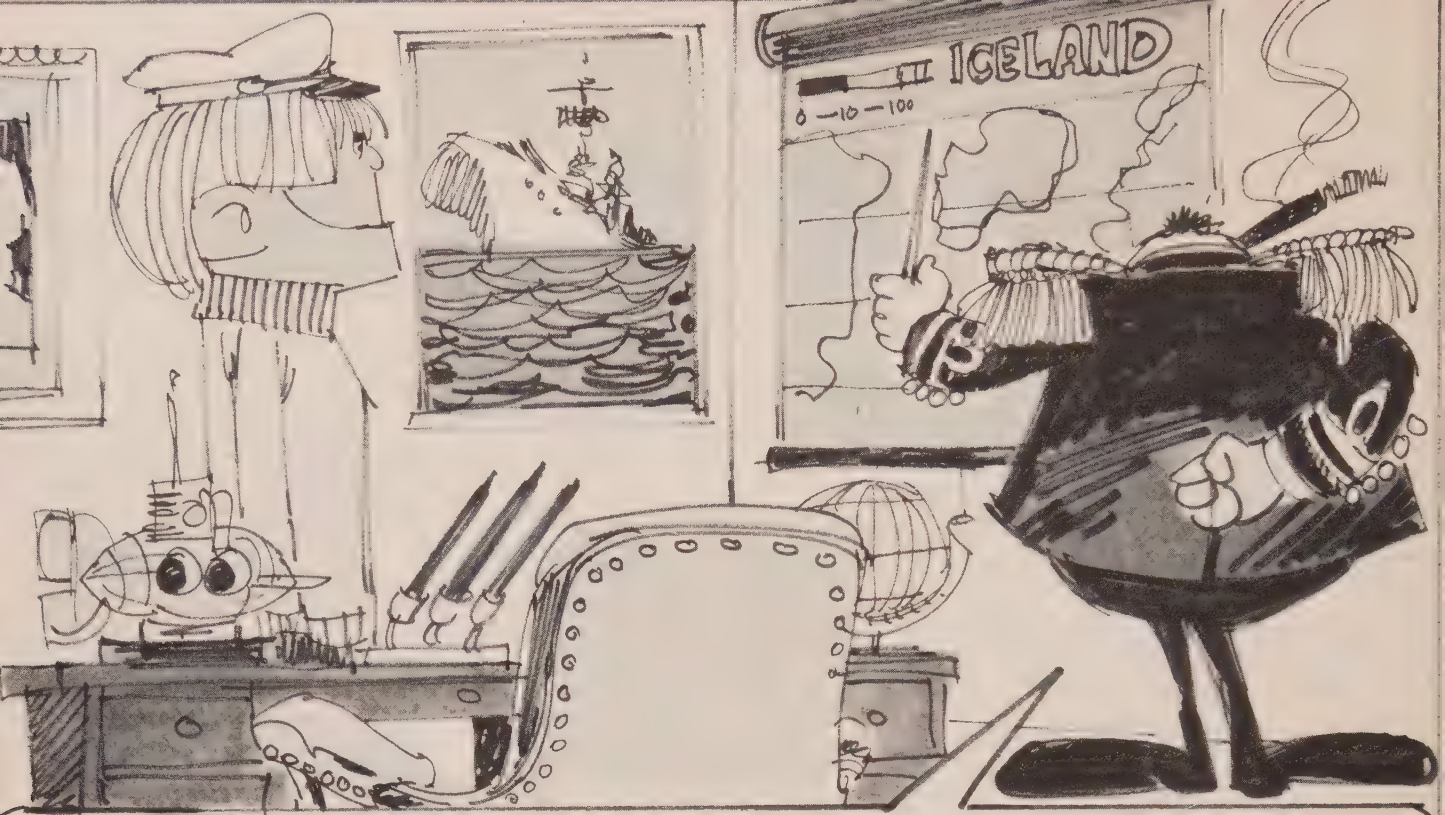


3. Read the temperature to the nearest $\frac{1}{2}$ degree after the mercury height remains steady. This reading is 86.5°C.





SAFETY SUGGESTION: Mercury from broken thermometers is a laboratory hazard. If you break a thermometer, have the teacher show you how to dispose of the spilled mercury.



Commander Seemore, I want you to explore the area around Iceland. Pay special attention to the volcanic heat and the boiling point of the ocean. Give me a full reading of the endothermic and exothermic changes. Then head north for Project DEEP FREEZE.

Water—No Quick-Change Artist

The solid, liquid, and gas forms of matter are known as PHASES. The solid phase of water is called ice; the liquid phase is called water; and the gaseous phase of water is called steam. During a change in phase, heat energy is taken in or released by the matter changing. *When energy is taken in, an* ENDOTHERMIC CHANGE *is taking place. When energy is released, an* EXOTHERMIC CHANGE *is taking place.*

Experiment 3.1

Equipment and Supplies per Lab Station

asbestos wire gauze
beaker—250 ml, burner
evaporating dish to fit on top of beaker (about 30 ml capacity), ring, ring stand
crushed ice to fill evaporating dish

What to do

1. Obtain and assemble the apparatus and crushed ice as shown in Fig. 7.
2. Slowly heat the beaker containing the ice. Record what you observe.

____

____

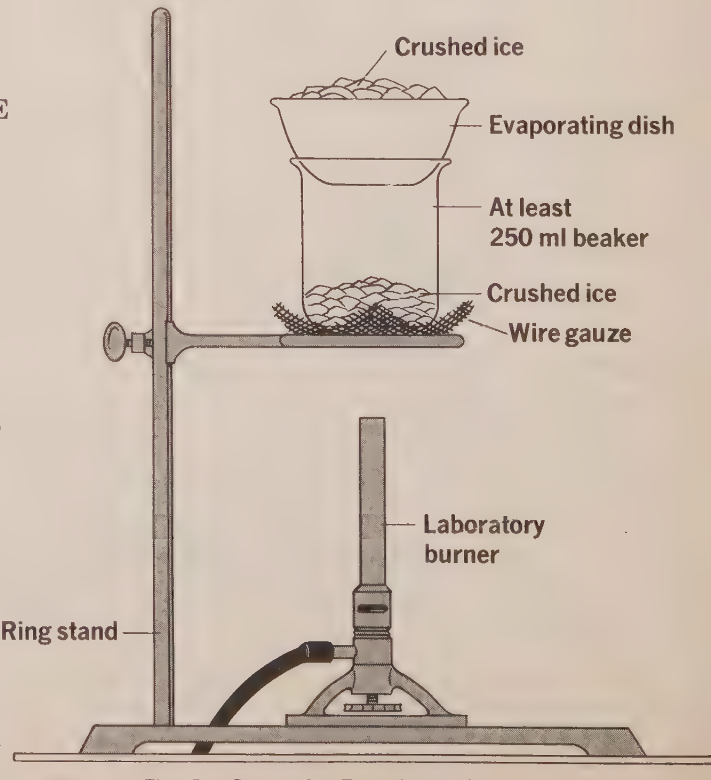


Fig. 7. Set up for Experiment 3.1.

3. Continue to heat the beaker slowly until boiling is started. Record what you observe about the contents of the beaker, the bottom of the evaporating dish, and the ice in the dish. Continue to heat the apparatus until all of the ice has melted in the evaporating dish.

4. Use a thermometer to take the temperature of the water in the dish. The reading is

   ______________°C.

5. In which phase changes did you observe heat taken in (endothermic); in which phase changes did you observe heat released (exothermic)? Write your answer in the Summary Chart.

SUMMARY CHART

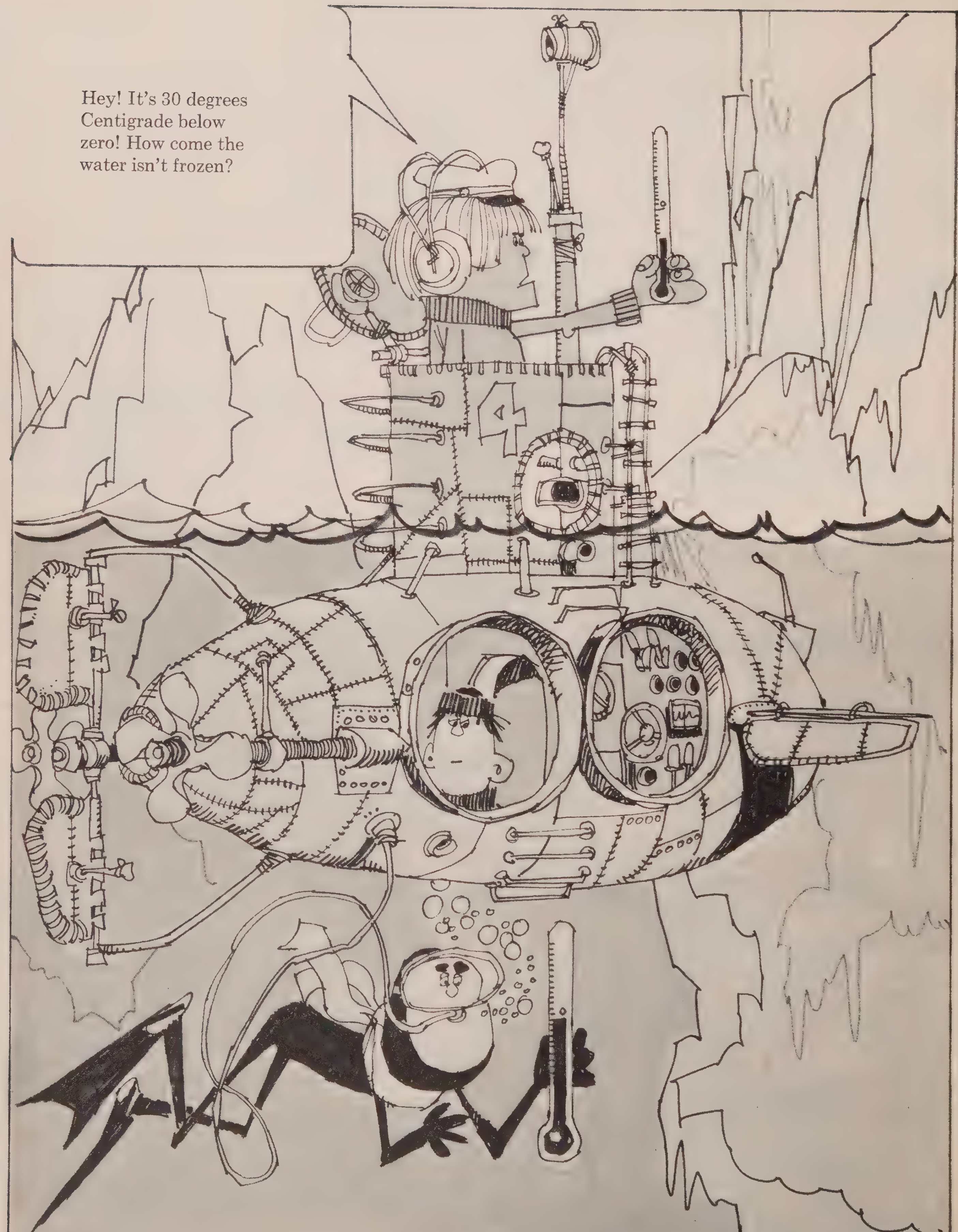
Observation	Phase Change		Endothermic Heat Taken In	Exothermic Heat Given Off
	from	to		

Extra for Experts

1. Why do you think steam is used in some heating systems?

2. Why does ice lower the temperature of matter in contact with it?

3. Why does heat applied to a substance change its temperature or make it change its phase?



Hey! It's 30 degrees Centigrade below zero! How come the water isn't frozen?
4

PROJECT DEEP FREEZE

Do you think that ice and water can exist at the same temperature? Can ice be colder than 0°C? Find the answers in PROJECT DEEP FREEZE!

Experiment 3.2

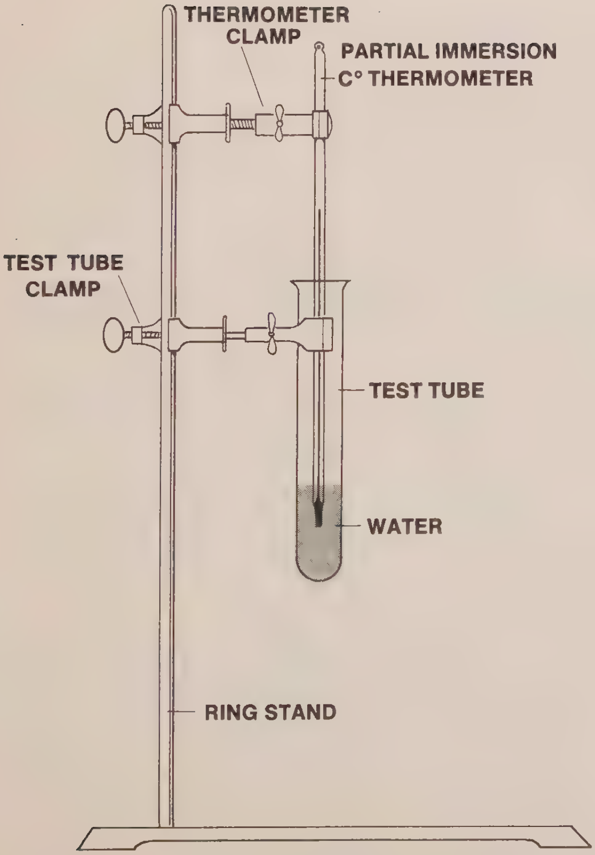


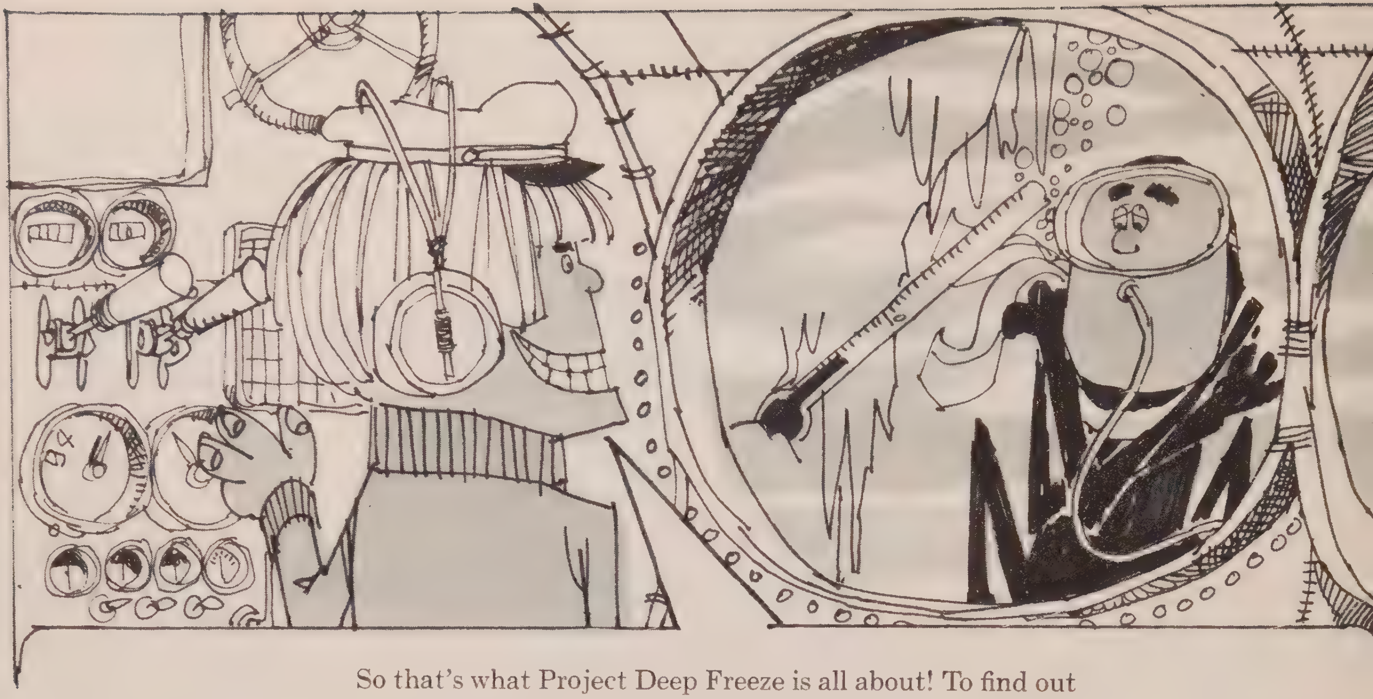
Fig. 8. Set up for Experiment 3.2 without ice.

Equipment and Supplies per Lab Station

beaker-400 ml or larger, or plastic-foam cup
(2) clamps—1 utility, 1 thermometer clamp
plastic spoon, ring stand, paper towels
thermometer with 1C° divisions to at least −5°C
distilled water, test tube, crushed ice
calcium chloride pellets or rock salt
coke or soda water or sugar in water

What to do

1. Set up the apparatus as shown in Fig. 8.
2. In a 400 ml (or larger) beaker mix a cupful of crushed ice with three teaspoonsful of calcium chloride pellets OR mix a cupful of crushed ice with 1/3 cupful of rock salt.
3. Put about a 1-inch layer of the ice mixture in the bottom of a plastic-foam cup.



So that's what Project Deep Freeze is all about! To find out if ice can be colder than zero degrees, and why the rest of the water doesn't freeze when it's so cold out.

4. Carefully lower the test tube and thermometer into the cup by loosening the clamps on the iron ring stand. (Fig. 9).

5. Pack ice mixture around the test tube so the water in the tube is below the surface of the ice.

6. Take the initial temperature and continue to read the temperature every 30 seconds (1/2 minute) until the water reaches −5°C. RECORD YOUR READINGS IN THE DATA CHART. Watch for the formation of ice and note the temperature. Ice is seen forming at ______________°C.

7. Raise the thermometer and test tube so the cup can be removed.
   (a) What has happened to the water?
   ____
   (b) At what temperature did this occur? ____________
   (c) Can ice be colder than 0°C? ______________
   What is the basis for your answer?
   ____
   ____

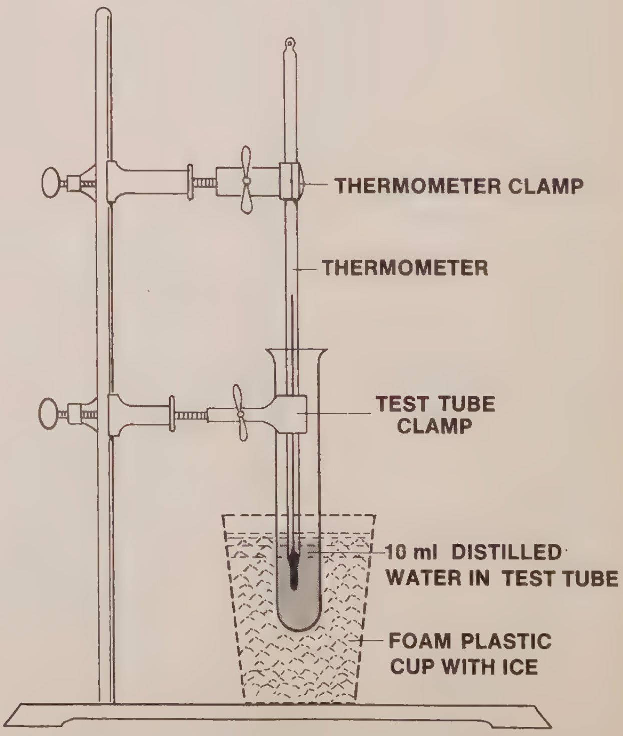


Fig. 9. Set up with foam cup with ice in it.

DATA CHART

Time Minutes	Temperature	Time	Temperature	Time	Temperature	Time	Temperature

Questions about the experiment

1. A graph can tell the story of what happened during OPERATION DEEP FREEZE.
   (a) Study Fig. 10 to see how to plot the point representing the temperature at a specified time.

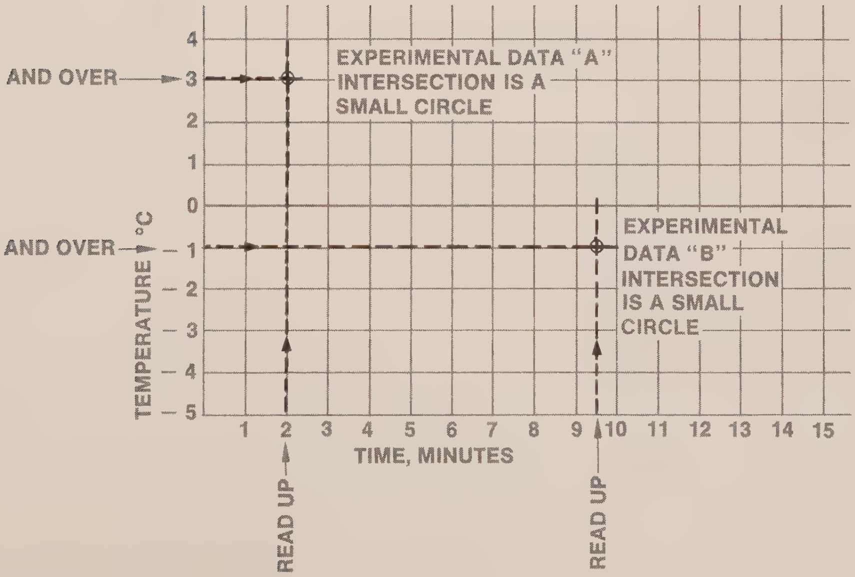
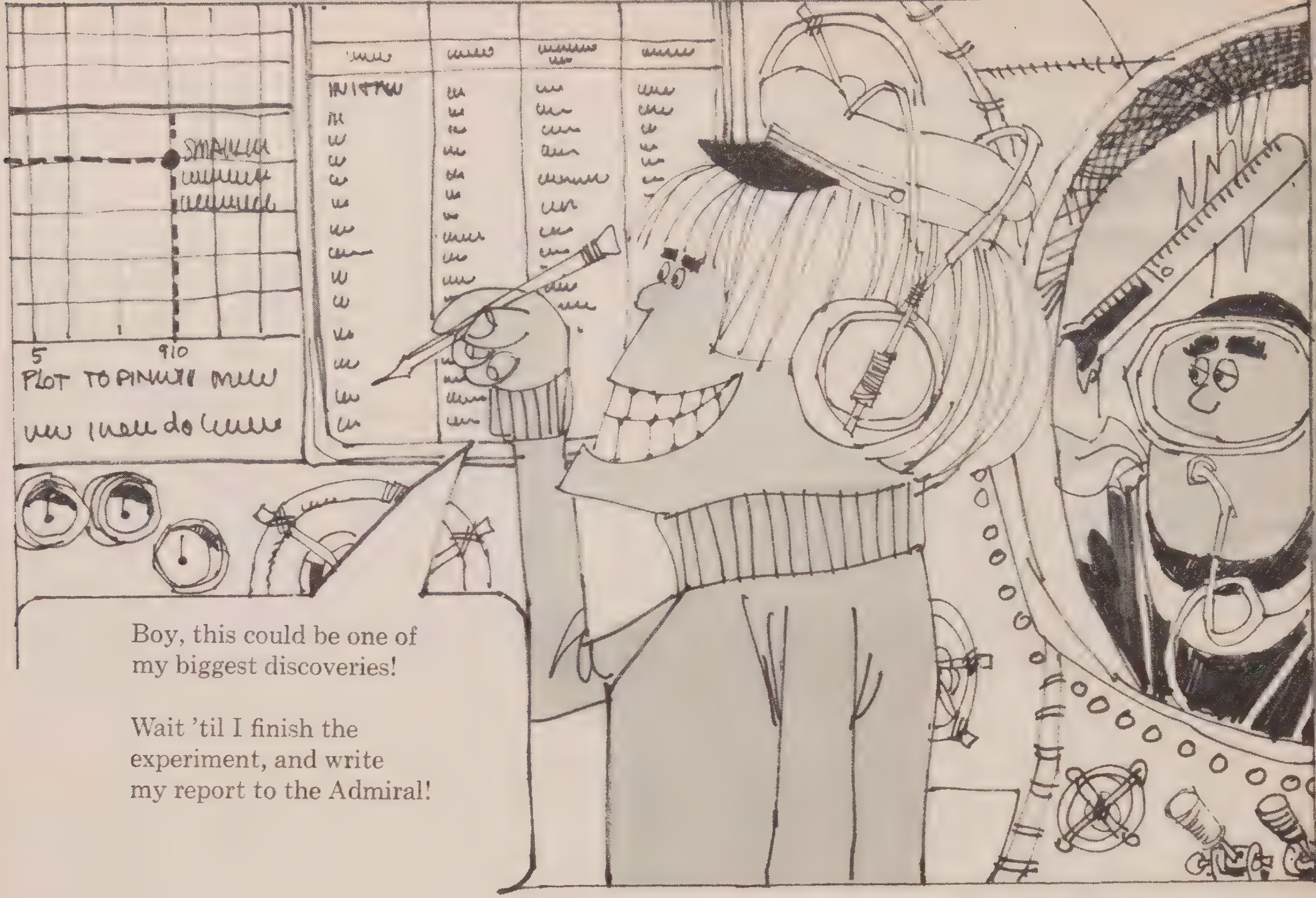


Fig. 10. A. Plot of the point representing 2 minutes at 3°C.
B. Plot of the point representing −1°C at 9½ minutes.


Boy, this could be one of my biggest discoveries!
Wait 'til I finish the experiment, and write my report to the Admiral!

(b) Now plot the data from the Data Chart on the graph grid below.

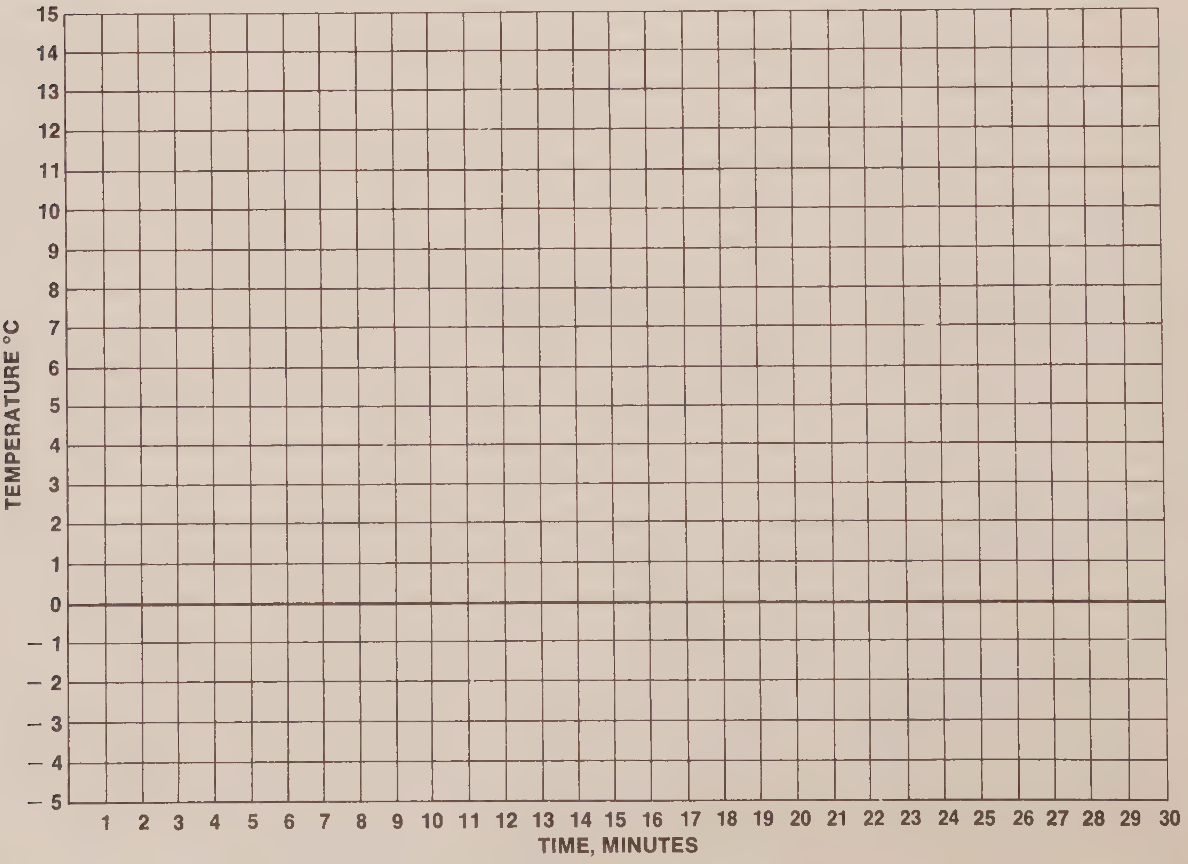


Fig. 11.

(c) Look for a general path to draw a smooth line through the circles representing the points. Try not to miss more than 2 or 3 dots.

(d) Mark the section of the graph where liquid water is cooling.

(e) On the plateau (a straight horizontal line) indicate that the water is freezing liquid → solid

(1) What does the rest of the graph tell us?

____

(2) What title should we put on the graph?

____

2. Does the temperature of the ice remain at 0°C when it melts? ____________ You might like to check your answer. Immerse a thermometer in crushed ice contained in a plastic-foam cup. Watch the thermometer as the ice melts.

3. When ice melts, is it an exothermic or endothermic change? ____________ What kind of change would the freezing of water be, exothermic or endothermic? ____________ What makes you think this?

____

____

4. Put an ice cube in a beaker of ice water, then decide which has the greater density, ice or ice water (0°C).

____________ has the greatest density.

How do you know this?

____

Extra for Experts

1. If you had started with steam instead of water and finally ended with ice, how would the graph differ from the one you made?

____

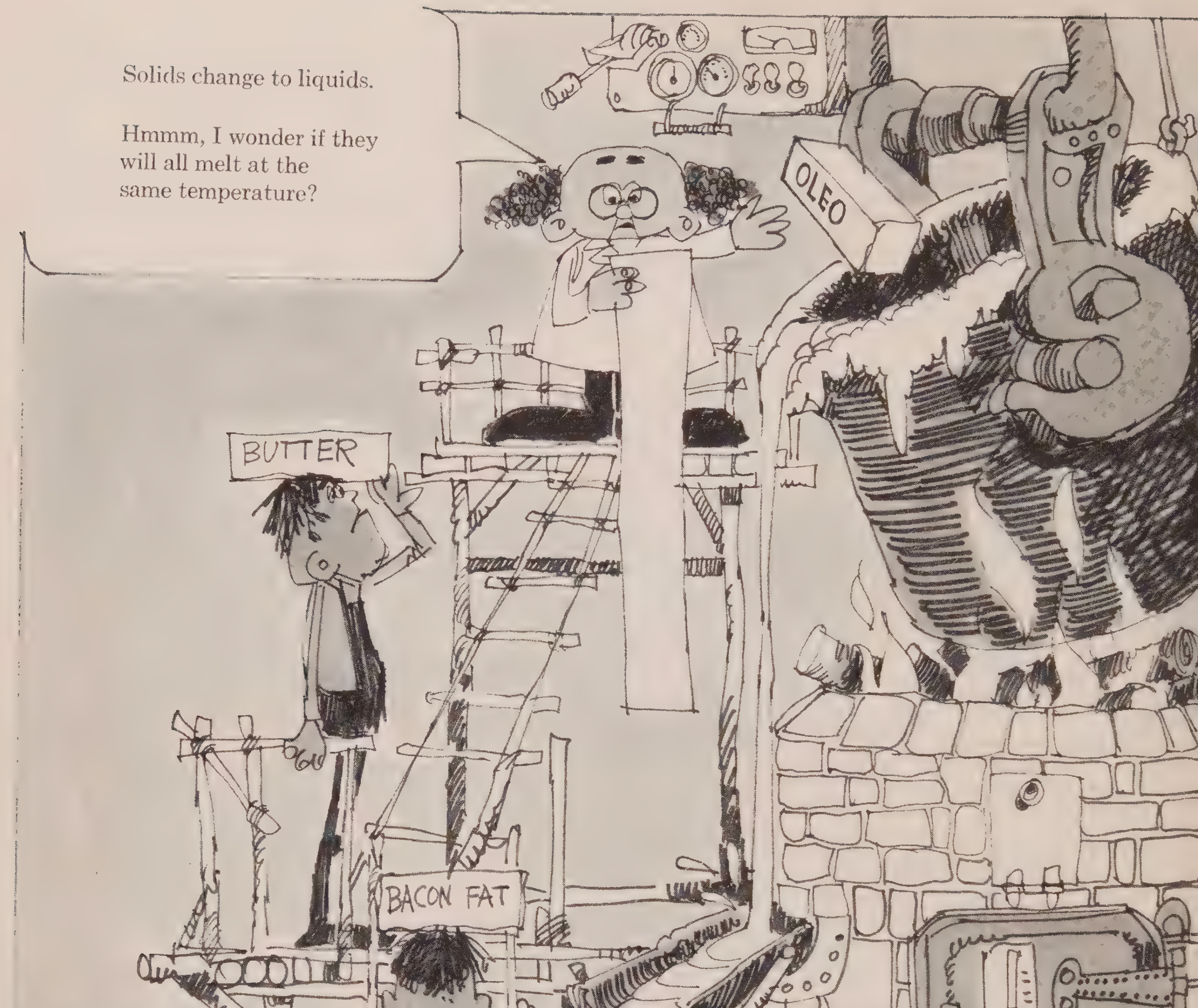
____

____

2. Will tap water, a sugar solution, or coke freeze at 0°C? Repeat the procedures using one of these liquids instead of distilled water to check your answer.

____

____


Solids change to liquids.

Hmmm, I wonder if they will all melt at the same temperature?
OLEO
BUTTER
BACON FAT
PIG IRON
CANDLE WAX

A HIGH CAL EXPERIMENT–MELTING FATS

Solid materials, when heated, may change to liquids. Do they all change at the same temperature?

In this experiment you may be able to answer this question for a few familiar materials.

Experiment 3.2–A

Extension

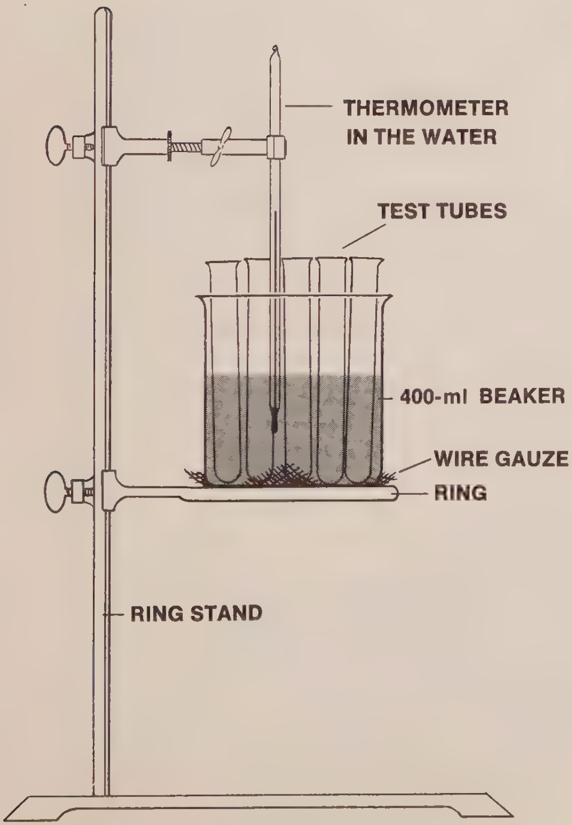


Fig. 12. Set up for Experiment 3.2-A.

Equipment and Supplies per Lab Station

asbestos wire gauze, beaker–400 ml
clamp–utility or thermometer type
ring stand with ring, burner
(5) test tubes, pyrex
test tube clamp or holder
thermometer, water
¼ teaspoonful of each:
bacon fat; butter; candle wax
oleomargarine; stearic acid; other assorted low-melting point solids assigned

What to do

Two students work together: One watches and records temperature readings; the other watches changes in the samples.

1. Obtain about ¼ teaspoonful of each of the following materials.
   (a) bacon fat
   (b) butter
   (c) candle wax
   (d) oleomargarine
   (e) stearic acid

2. Number five test tubes: 1, 2, 3, 4, 5.
3. Transfer the samples to the numbered test tubes. Follow the order given in the list in step 1.
4. Half fill your 400-ml beaker with tap water. Place all five of the test tubes with the samples in the beaker.
5. Arrange the apparatus as shown in Fig. 12. Clamp the thermometer to the ring stand so that the bulb is in the water in the beaker.
6. What is the temperature of the water in the beaker? ________°C.
7. Light your burner. Adjust the flame to be moderately hot. Place the burner under the beaker so that the tip of the flame is about one inch below the wire gauze.
8. Watch the samples for the first sign of melting. (Don't wait until all of the sample has melted!) Immediately take the temperature of the water. While one partner records the temperature in the Chart below, the other partner should use the test tube clamp to remove the test tube of melting substance from the beaker. Continue until all of the test tubes have been removed.

DATA CHART

TABLE OF MELTING TEMPERATURES

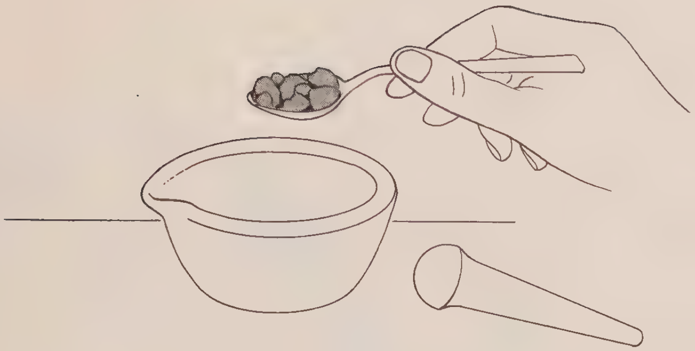
Test Tube	Sample	Melting Temperature °C	Comment
1			
2			
3			
4			
5			

Your conclusions

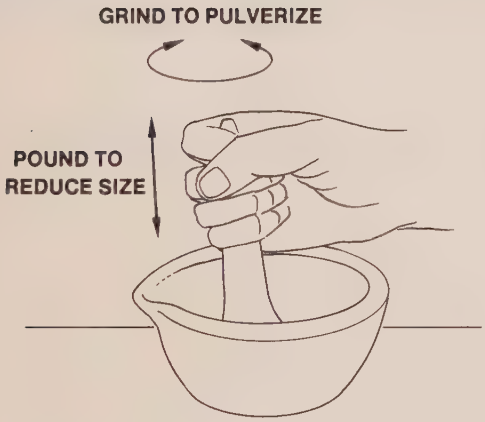
1. Did all of your samples melt at the same temperature? ________________________
2. Do some materials need more heat than others before they melt? ________________________

HOW TO

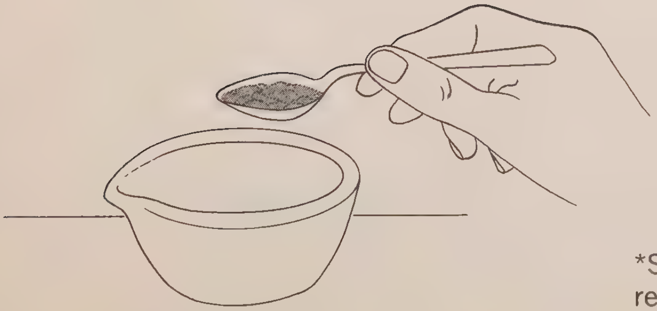
use a mortar and pestle



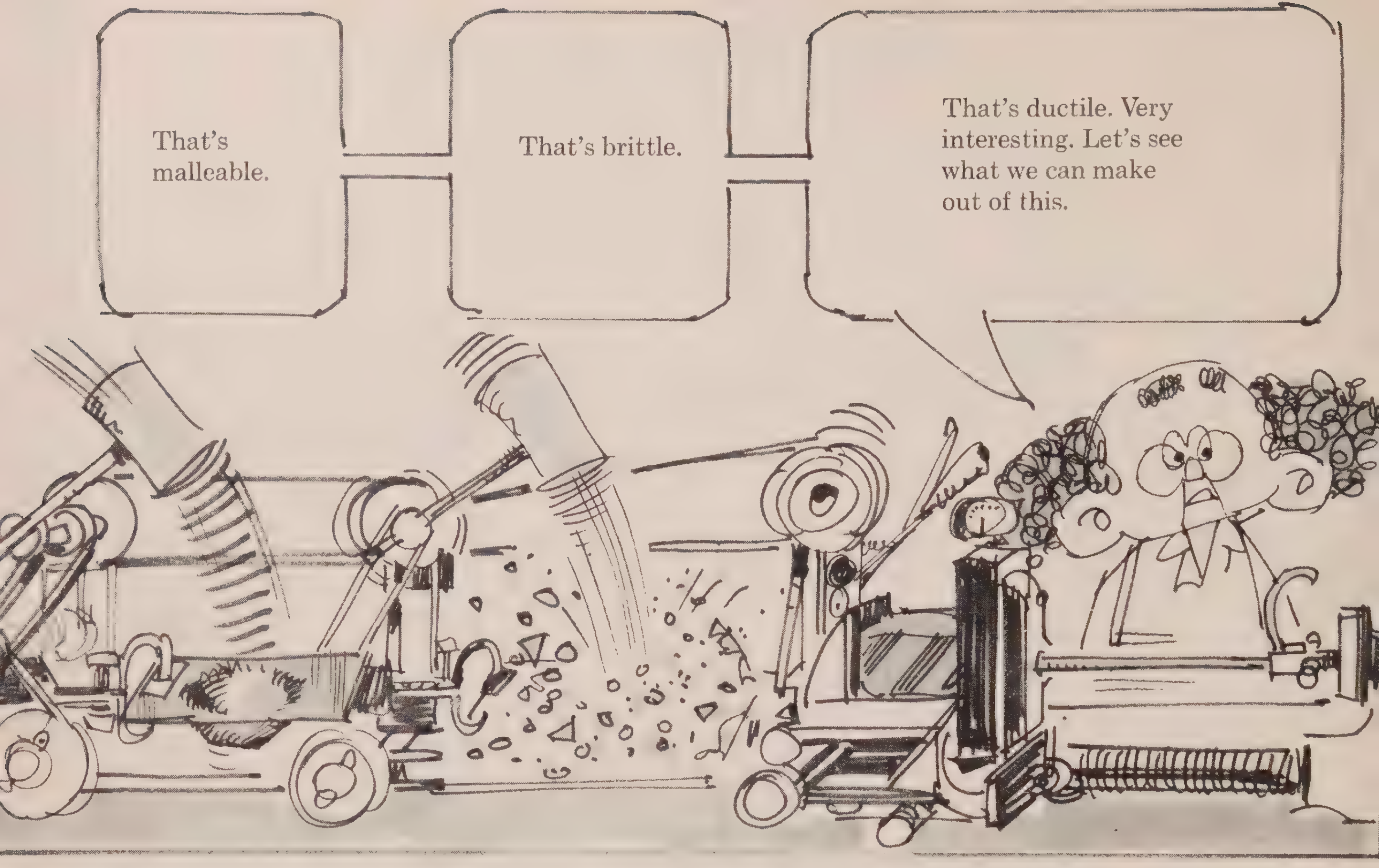
Place material to be broken up into mortar.



Pound with pestle.



*Safety reminder: Do not blow into mortar to remove any remaining powder. Dust may get into your eyes.


That's malleable.
That's brittle.
That's ductile. Very interesting. Let's see what we can make out of this.

TELLING METALS FROM NONMETALS

Chemical substances, like people, have individual characteristics. Chemists call these characteristics PHYSICAL PROPERTIES.

Examine the properties of a sample of each of the substances provided. *PHYSICAL PROPERTIES are those characteristics we learn from our physical senses.* Use the likenesses and differences in properties to arrange the substances into two or more groups.

Experiment 3.3

Equipment and Supplies per Lab Station

conductivity apparatus (at least 1 per class)
containers for carrying sample to workbench
marking pencils, mortar and pestle
carbon (wood charcoal), copper
lead pieces, sulfur (roll), other samples

What to do

1. Obtain a small sample of each of the substances. Be sure to label each.
2. Record the color of each sample in the Chart of Properties.

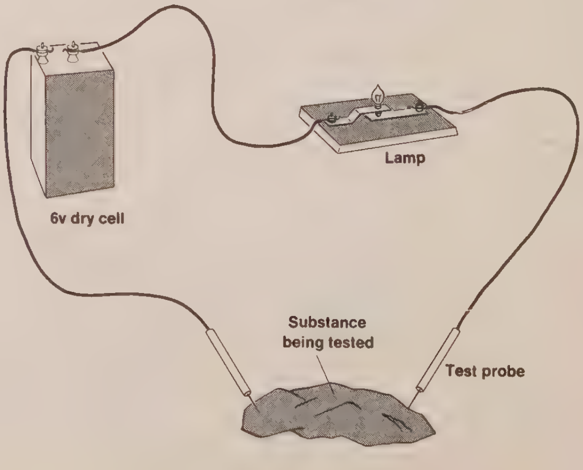


Fig. 13. Conductivity apparatus.

3. Place the samples, one at a time, in your mortar and pound each with the pestle. Wipe the mortar between samples.
    (a) *If the substance can be pounded into a shape, without crumbling, record the substance as MALLEABLE.*
    (b) *If the substance crumbles into small particles as it is pounded, record it as BRITTLE.*
4. Take your samples to the conductivity apparatus and test the ability of each sample to conduct electrical current.
5. *DUCTILITY is the ability of a substance to be drawn into a wire.* Inspect the stock bottles of substances to see which contain the substance in wire form. Add the data to the Chart.

CHART OF PROPERTIES

Property	Copper	Lead	Carbon	Sulfur	Other
1. Color shiny dull					
2. Malleable					
3. Brittle					
4. Electrical Conductor					
5. Ductility					

6. Study your Chart of Properties. Arrange the substances in two or more groups according to likenesses. For each group list the properties which are alike.
7. Read pages 32–34 in the text to find out the properties of metals and nonmetals. Then on the basis of your observations list which samples are metals and which are nonmetals.

Metals ______________________________

____

Nonmetals ______________________________

____

Questions to answer

1. Metals are good conductors of electricity. What other physical properties do metals have?

____

2. Nonmetals may have physical properties which are just the opposite of metals. List the properties which nonmetals have.

____

What is bright yellow, brittle and has a boiling point of 444.6 degrees centigrade?
Hmmm, let me see? I know! A rabbit that's been scared stiff and boiling mad?

Exercise

what is it?

In Experiment 3.3 you found that metals had some general physical properties. Nonmetals had a different set of physical properties. In addition to the general properties, each metal or nonmetal has some special properties of its own.

The clues in some of the puzzles are just general properties, so you may be able only to classify the substance as a metal or a nonmetal. Some clues give the special properties. Information in the Reference Table in the APPENDIX will help you use these special properties to identify the substances by name.

1. Generally it is recognized by a bright yellow color. It is brittle and has a boiling point of 444.6°C.

   ____

2. This foil has a luster and can conduct electricity.

   ____

3. A dull surface hides the luster of this solid which has a density of 11.34 g/cm³. The great density makes it possible for the solid to hold a fishline under fresh water.

   ____

4. You can't see this colorless gas which is slightly denser than hydrogen. It is used to fill balloons.

____

5. This yellow, malleable substance is used to make rings. Its density is 19.3 g/cm^3.

____

6. Sometimes you find this substance as gray crystals. At higher temperatures you can see it as a violet gas. It is often used in making an antiseptic to treat cuts.

____

7. There is a rosy luster to this electrical conductor's appearance. It melts at approximately 272°C.

____

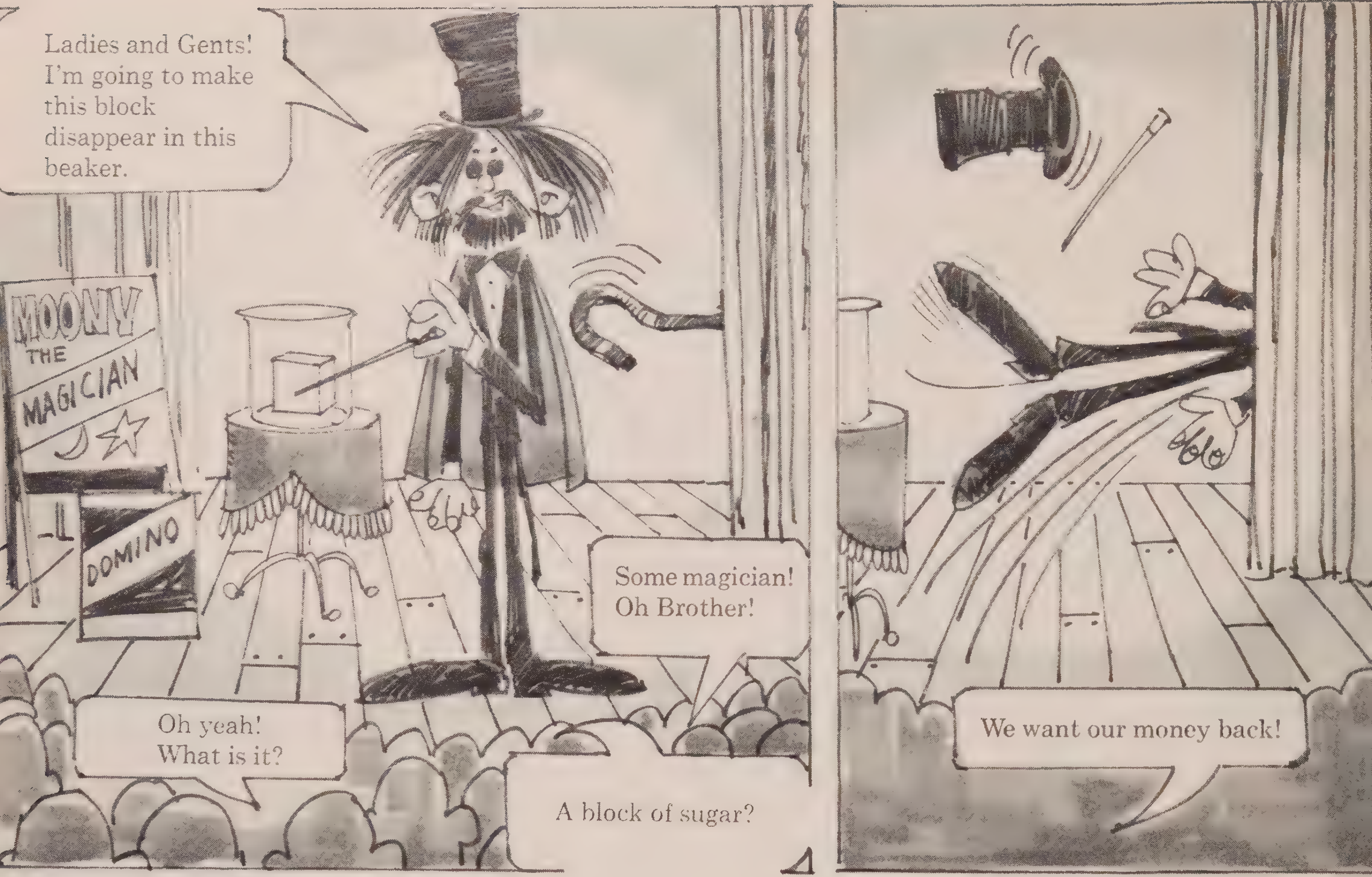
8. Scientist X has discovered a silvery white solid which is lighter in density than aluminum.

____

9. This material is a dark red liquid at room temperature. It does not conduct electricity.

____

10. This element is dull and brittle. It is a non-conductor of electricity. _______________


Ladies and Gents! I'm going to make this block disappear in this beaker.
MOONY THE MAGICIAN
DOMINO
Some magician! Oh Brother!
Oh yeah! What is it?
A block of sugar?
We want our money back!

a chemical disappearance

Did you ever think the disappearance of sugar and table salt in water was mysterious? Chemists have found clues to the mystery. You will learn more about their ideas on such chemical disappearances later in this course.

Follow the clues to find some answers to the two questions:

1. Do other solid chemicals disappear in water the way sugar and table salt do?
2. May color be a clue to where a chemical goes in water?

Experiment 3.4

Equipment and Supplies per Lab Station

balance, marking pencil, metric ruler
filter paper or container for weighing
stirring rod, (6) test tubes, test tube rack.

Two gram samples of each of the following:
calcium hydroxide, cobalt (II) nitrate
iron filings, iron (II) sulfate
nickel (II) sulfate, sodium chloride (table salt)

Some useful terms

1. *SOLUTION is a mixture of two or more substances; the mixture is alike throughout. Its composition may be varied within definite limits.*
2. *SOLUTE – the substance being dissolved*
3. *SOLVENT – the substance doing the dissolving*
4. *SOLUBLE – capable of being dissolved*

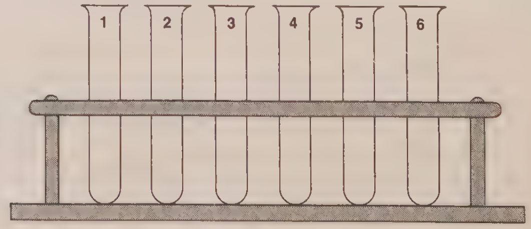


Fig. 14. Numbered test tubes.

What to do

1. Number six test tubes. Line them up as illustrated.
2. To each test tube add 2 grams of one of the six chemicals. Follow the order in the table below. Record the color of each chemical in the Chart.
3. Shake the dry chemical well down into the bottom of its test tube. Mark the top level. Measure in millimeters the height of this level above the bottom of the test tube as shown in Figure 15. In the Chart, record this measurement for each chemical.
4. Add 10 ml of water to each test tube. Stopper and shake each test tube very hard for about half a minute. Let stand for five to eight minutes. Mark, measure, and record the level in millimeters of any undissolved (not disappeared) chemical.
5. Complete the Chart by recording the color of the water plus chemical in each test tube.
6. Study the Data Chart.

(a) What differences did you observe in the dissolving of the six chemicals tested?

____

____

(b) Where does the disappearing solid go in the water? The color will give you a clue.

____

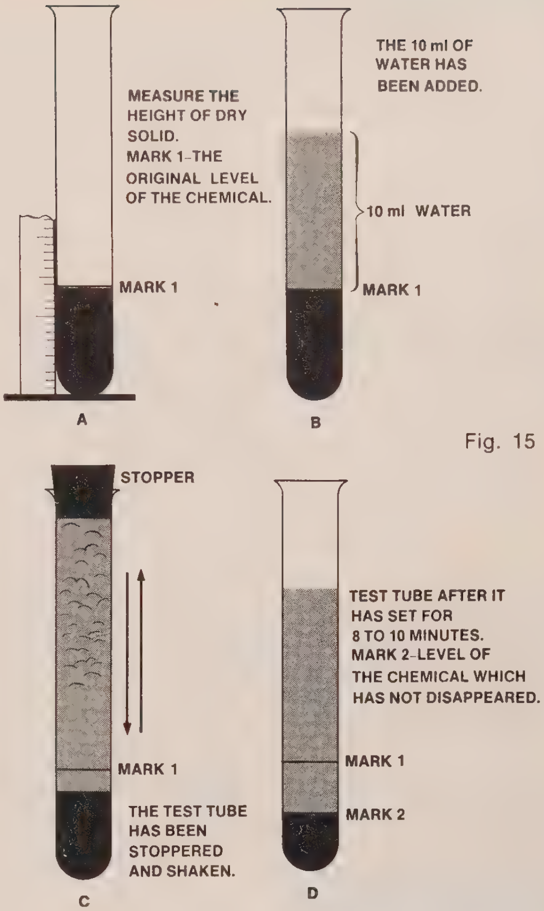

MEASURE THE HEIGHT OF DRY SOLID. MARK 1–THE ORIGINAL LEVEL OF THE CHEMICAL.
MARK 1
A
THE 10 ml OF WATER HAS BEEN ADDED.
10 ml WATER
MARK 1
B
STOPPER
MARK 1
THE TEST TUBE HAS BEEN STOPPERED AND SHAKEN.
C
TEST TUBE AFTER IT HAS SET FOR 8 TO 10 MINUTES. MARK 2–LEVEL OF THE CHEMICAL WHICH HAS NOT DISAPPEARED.
MARK 1
MARK 2
D

Fig. 15

DATA CHART

OBSERVATIONS					
Test Tube No.	Chemical Tested	Color of dry Chemical	Height of dry Chem. (mm)	Height of Undissolved Chem. (mm)	Color of Water plus Chemical
1	Iron filings				
2	Table Salt				
3	Iron (II) Sulfate				
4	Calcium Hydroxide				
5	Nickel Sulfate				
6	Cobalt Nitrate				

Questions about the experiment

1. Which of the chemicals tested in this experiment formed solutions?

   ____

2. In Test Tube No. 3, what was
   (a) the solute? ______________
   (b) the solvent? ______________

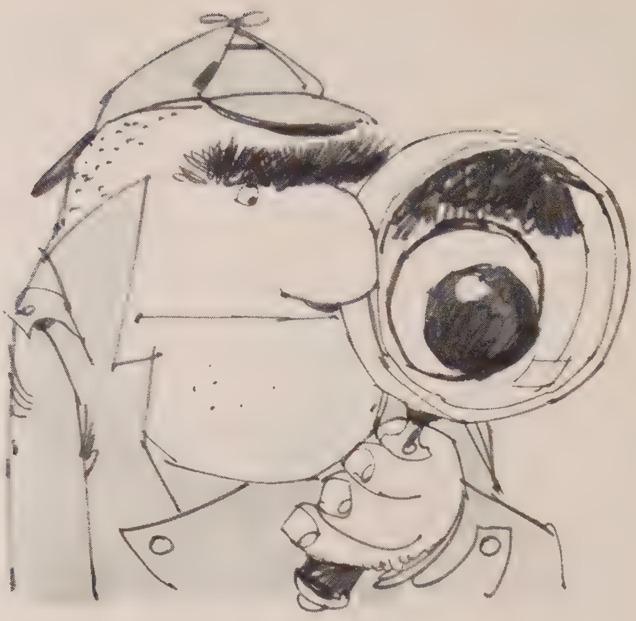
3. Of the chemicals tested, which ones, under the conditions of the experiment, showed
   (a) very high solubility in water?

   ____

   (b) very low solubility in water?

   ____

Chapter 4



some detective work

the look-alikes and the unlook-alikes

Matter may be classified as an element, mixture, or compound. Among the samples given you, are the elements sulfur and copper. The other samples are either mixtures or compounds.

Properties are the "chemical fingerprints" to be used in making identifications. Identify which samples are the mixtures and which are the compounds.

Experiment 4.1

Equipment and Supplies per Lab Station

containers for carrying samples to workbench
samples from six sources

Some clues

1. Elements lose their properties when they combine to form a compound which has its own new set of properties.
2. The substances in a mixture can be identified by their respective properties.
3. Contents of the jars are only combinations of the elements Cu and S, their compounds, or mixtures.

Making the identification

1. Look at the samples of the elements: copper and sulfur. How can you recognize each one?

copper ______________________________

sulfur ______________________________

2. Carefully observe each of the samples marked X, Y, Z, and EZ. Fill in the Data Chart for each sample.

DATA CHART

Substance	Appearance	How many substances observed?
X		
Y		
Z		
EZ		

3. Study the clues and the Data Chart. Then make an identification:
   (a) Which is the mixture of the element form of copper and sulfur?

   ____

   ____

   (b) Which one is the compound of copper sulfide?

   ____

4. Check your identification of the samples. Pull off the tape on each container. Read the label which you can see now. Correct your answers if necessary.

5. Replace the tape so the list of contents can not be seen by the next student using the samples.

Questions to answer

1. Must a mixture always be made up of just two elements?

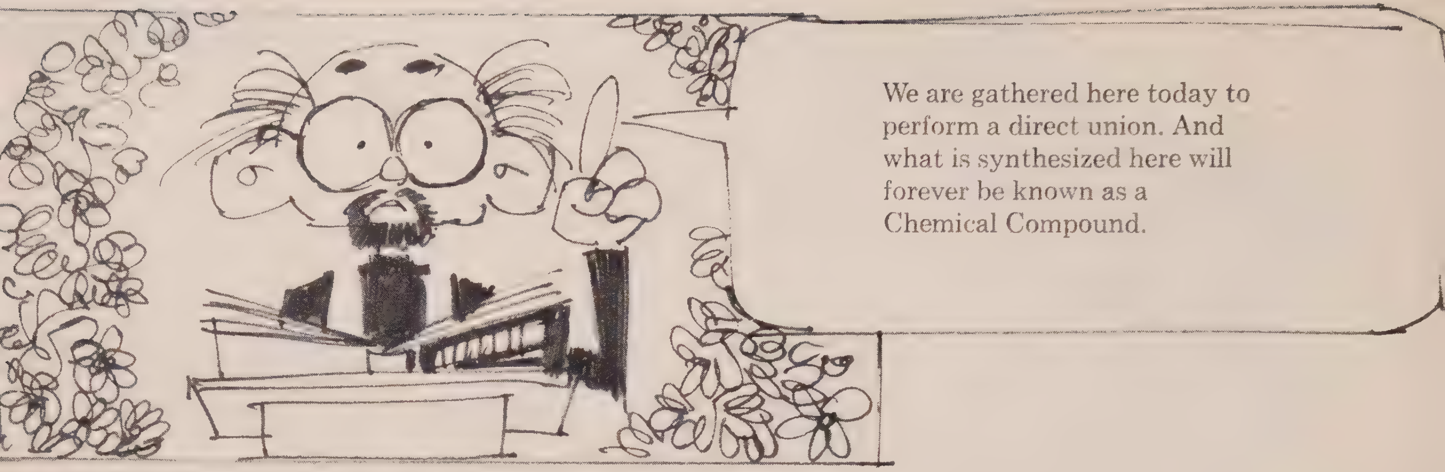
   ____

2. Can a compound be formed from just one element?

   ____

   ____

   ____


We are gathered here today to perform a direct union. And what is synthesized here will forever be known as a Chemical Compound.

a chemical merger

Two or more substances may combine to form a compound. Chemists call this type of chemical action a SYNTHESIS or a DIRECT UNION.

In this experiment you will heat copper in air and find out whether you have synthesized a compound.

Experiment 4.2

Equipment and Supplies per Lab Station

asbestos pad, asbestos wire gauze
balance, burner, evaporating dish
ring, ring stand, test tube, tongs
copper wire (22 gauge) – 70 cm or 190 cm of 26 gauge

What to do

1. Find the mass of a clean, dry evaporating dish. Record the mass in the Data Chart.
2. Obtain a length of fine copper wire.
3. Coil the wire loosely around a test tube.
4. Carefully remove the coil and set it in the evaporating dish.
5. Find the mass of the dish and the copper coil. Record the combined mass in the Data Chart 1.

DATA CHART 1

Mass of the dish and copper ______________ g

Mass of the empty dish ______________ g

Mass of copper wire (subtract above) ______________ g

6. Assemble the apparatus as shown in Fig. 16. Gently heat the evaporating dish for 5 to 10 minutes.
7. Use tongs to transfer the dish and its contents to a wire gauze on the desk. See Fig. 17. Cool the dish and its contents to room temperature.
8. While the evaporating dish is cooling, examine *but do not touch* the contents. Describe any changes you see.

____

____

9. Find the mass of the dish and its contents. Record the mass of the dish and contents (after heating) in Data Chart 2.

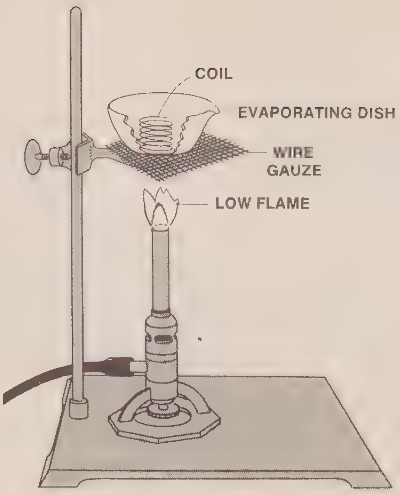


Fig. 16. Set up for Experiment 4.2.

DATA CHART 2

Mass of dish and its contents after heating	______________ g
Mass of dish and its contents before heating (Line 1 in Data Chart 1)	______________ g
Difference in mass (subtract the masses)	______________ g

Questions about the experiment

1. Did your data show a change in mass of the copper?

____

How can you account for this change? ____________

2. What observations show that a chemical compound may have been formed?

(a) ____

(b) ____

3. If a copper compound has been formed, with what may the copper have united? ____________________

What would be the name of the compound? __________

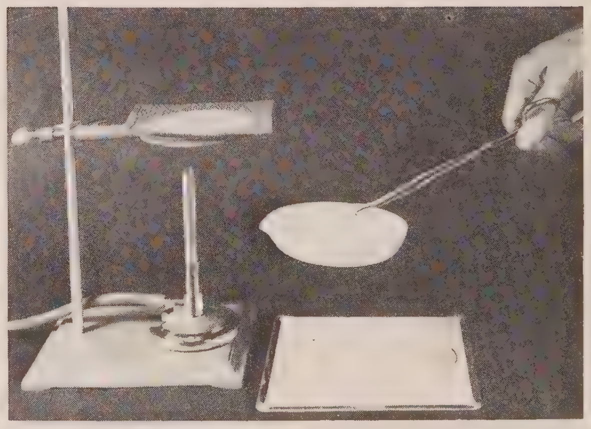


Fig. 17. Transfer evaporating dish to asbestos pad.

Extra for Experts

What is the mass of the possible copper compound you made in your experiment? ____________________

Writing Direct Union Word Equations

Exercise

PART I

Directions: Write the word equation for the reaction between each of the following elements:

1. sulfur and zinc →
2. aluminum and oxygen →
3. copper and chlorine →
4. magnesium and iodine →
5. bromine and sodium →
6. hydrogen and chlorine →
7. hydrogen and sulfur →
8. calcium and chlorine →
9. potassium and iodine →
10. oxygen and zinc →

PART II

Complete the following word equations by filling in the blanks:

1. ______________ + oxygen → zinc oxide
2. potassium + ______________ → potassium iodide
3. ______________ + oxygen → carbon dioxide
4. ______________ + ______________ → sulfur dioxide
5. ______________ + ______________ → calcium iodide
6. ______________ + ______________ → sodium chloride
7. magnesium + ______________ → magnesium bromide
8. ______________ + ______________ → carbon disulfide
9. ______________ + ______________ → zinc sulfide
10. ______________ + ______________ → water

Oh, Rex! Rex! wherefor art thou? Come and do thy Chemistry experiment!
What's in a chemical? That which we call a NaCl by any other word would taste as salty; So salt would react, were it called NaCl, in the same way with another chemical. Oh, what puzzles me and thee is not the name, but the chemical properties!

A Chemical Alias Or What's In A Name?

All samples of the same chemical must react in the same way with another chemical. You can solve the puzzle below by comparing the reactions of known substances with the reactions of your unknowns.
The puzzle: What chemical is each of the household items provided by your teacher?
Rules of the puzzle: You must identify each substance on the basis of chemical not physical properties. The reactions with other chemicals indicate chemical properties.

Experiment 4.3

Equipment and Supplies per Lab Station

beaker, dropper, (2) spoons
(6) test tubes, test tube rack
Epsom salts, solid, distilled water
household ammonia
litmus test paper—red and blue
sodium chloride crystals, white vinegar
2–3 ml of each of the following solutions:
acetic acid, ammonia water, barium chloride
cobalt (II) nitrate, magnesium sulfate
potassium bromide, sodium chloride

What to do

1. If the substance given you is not a solution, dissolve about a teaspoonful of the solid in 100 ml of distilled water.
2. Test the unknown with red and blue litmus papers. Record your observations in Data Chart 1.
3. Put 10 drops of the unknown solution into a small test tube. Add 5–10 drops of cobalt nitrate solution. Record any observations in Data Chart 1.
4. Put 10 drops of the unknown solution into another small test tube. Add 10 drops of barium chloride solution. Record any observations in Data Chart 1.
5. Repeat steps 2–4 with each of the unknowns provided by your teacher. Write your observations in Data Chart 1.
6. Repeat steps 2–4 with each of the known chemicals provided by your teacher. Record these observations in Data Chart 2.

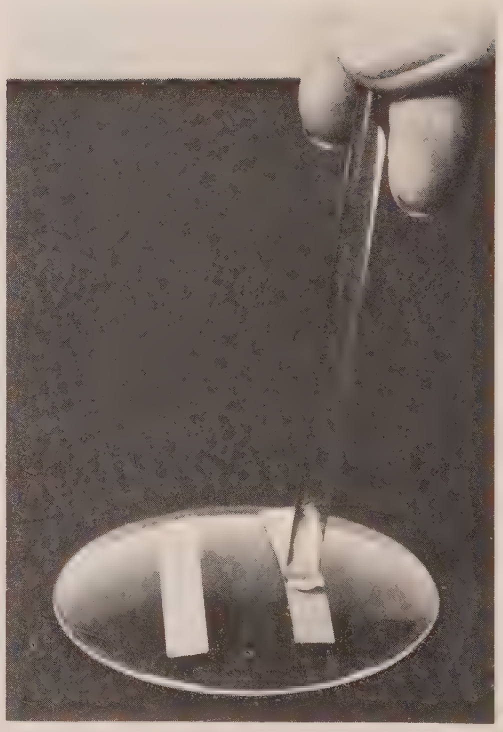


Fig. 18. Using litmus paper.

DATA CHART 1

Unknown Substance Given You	Reaction with			
	Red litmus	Blue litmus	$Co(NO_3)_2$	$BaCl_2$
1.				
2.				
3.				
4.				

DATA CHART 2

Known Substance	Reaction with			
	Red litmus	Blue litmus	$Co(NO_3)_2$	$BaCl_2$
NaCl				
CH_3COOH				
$MgSO_4 \cdot 7H_2O$				
KBr				
NH_4OH				

7. Study DATA CHARTS 1 and 2, then complete the Puzzle Answer CHART.

PUZZLE ANSWER CHART

Unknown Sample No.	Sample is identified as
1.	
2.	
3.	
4.	

How am I going to find out how many different things are in this mixture? Why, there could be thousands and even hundreds!

unscrambling–recovering the parts

Can you recover the original substances used to make mixtures or compounds? Discover the answer to this question by doing this experiment.

Experiment 4.4

Equipment and Supplies per Lab Station

asbestos pad, beaker–250 ml
evaporating dish, filter paper, funnel
magnet, paper towels, plastic bag,
plastic spoon, ring stand with a ring
test tube, tongs, triangular pipe stem
wire gauze to be used as screen
mixture to be separated

Part I–Separation of a mixture

How many different kind of substances can you separate out of a given mixture?

What to do

1. Examine two spoonfuls of the mixture given you. How many substances do you think are in the mixture?
   ____
2. Sift the mixture through a screen held over a paper towel. How many different substances are left on the screen? ______________________
3. Use your tongs to remove the pieces left on the screen. Put those that look alike on the same sheet of paper or in the same test tube.
4. Put a plastic bag or a piece of towel around the end of a magnet. Touch the magnet to the material that passed through the screen. Lift the magnet and its wrapping.
5. Hold the magnet over a piece of paper or a beaker. Carefully remove the magnet from the bag or towel. The material clinging to the magnet will drop on the paper or in the beaker.

6. Repeat steps 4 and 5 until no more material clings to the magnet. Put all of the particles removed by the magnet on the same paper or in the same beaker.
7. Examine the material left unattracted to the magnet. How many substances do you have left?

   ____

8. Pour 20 ml of water into a beaker. Add the material left unattracted to the magnet. Stir for two or three minutes. Did anything seem to dissolve in the water?

   ____

9. Read the HOW TO FILTER page, then set up a filtering apparatus. Filter the substances in the beaker.
10. Examine the RESIDUE (*what's left on the filter paper*). Place the filter paper and the residue in a beaker or on a paper towel.
11. Set up an evaporating apparatus. Evaporate the FILTRATE *(the liquid that passed through the filter paper).*
12. Examine any substance found in the dry evaporating dish. If there is anything in the dish, from where must it have come?

    ____

13. Compare the amount of each substance you have obtained with the amounts obtained by other students. Did all of the mixtures contain the same amount of the various substances?

    ____

    What does this tell you about mixtures?

    ____

14. Discard any wet or damp substances. Mix the dry pieces of your mixture together again. Put them back into the jar for use by other classes.

Questions to answer

1. Did you guess the right number of different kinds of substances in step 1? ______________

   Why were or weren't you able to do this?

   ____

2. How can a mixture of oil and water be separated?

   ____

3. If you were blindfolded, could you separate a mixture of gumdrops, peanuts, and caramels?

   ____

   How would you do it? ______________________

4. Could you separate a mixture of table salt and sugar the way you separated the mixture in this experiment? ______________________

   What are the reasons for your answer? __________

   ____

   ____

Part II Separation of a Compound

How many different substances can you separate out of the compound, table sugar?

Equipment and Supplies per Lab Station

asbestos wire gauze, burner
crucible tongs, evaporating dish, plastic spoon, ring stand with ring
cobalt chloride test papers
table suger (sucrose), distilled water

What to do

1. Examine some table sugar. How many substances do you think are in the sugar? ____________
2. Describe the appearance of the sugar.

   ____

3. Place a spoonful of sugar in an evaporating dish.
4. Put a few drops of water on a strip of cobalt chloride test paper. What happens to the color of this test paper when water is present? ____________
5. Using tongs, hold a dry strip of cobalt chloride test paper above the sugar in the evaporating dish. Touch the paper to the crystals of sugar. Does the test paper show the presence of water in the sugar?

   ____

6. Assemble a ring stand with a ring, an asbestos wire gauze, and a burner. Set the evaporating dish on the gauze.
7. Gently heat the sugar until it melts and another noticeable change occurs. Write down all you observe.

   ____

   ____

8. Use tongs to hold a dry piece of cobalt chloride test paper above the hot, changing sugar. Does the color of the test paper change?

   ____________ If so, what substance is being released from the sugar?

   ____

9. When no further changes are observed, turn off the burner, and let the dish and its contents cool. What does the substance left in the dish look like?

   ____

Something to do and answer

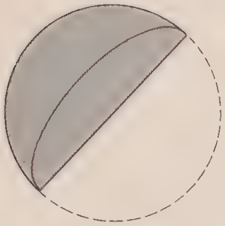
1. If you mix water with the residue in the dish, do you get sugar again? ____________
2. How does the separation of a mixture differ from the separation of a compound?

   ____

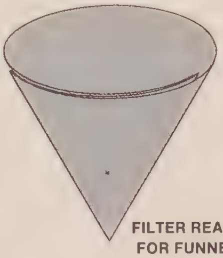
   ____

HOW TO

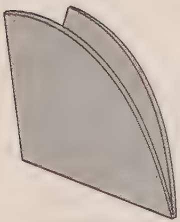
Use a Filter and Evaporate a Liquid



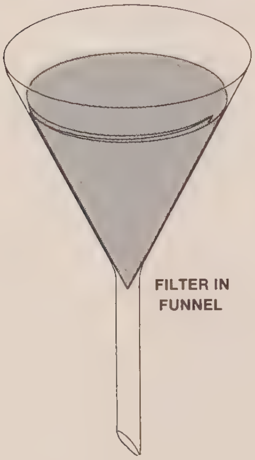
FILTER FOLDED ALONG DIAMETER



FILTER READY FOR FUNNEL

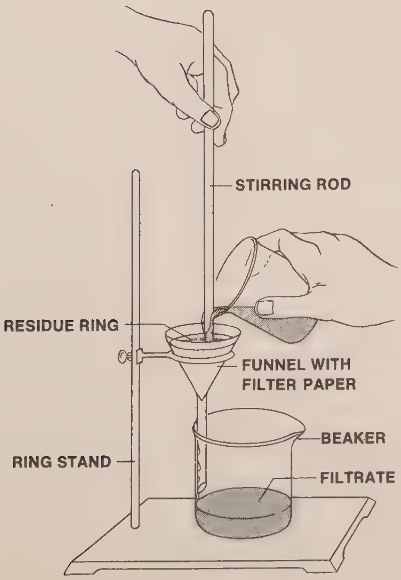


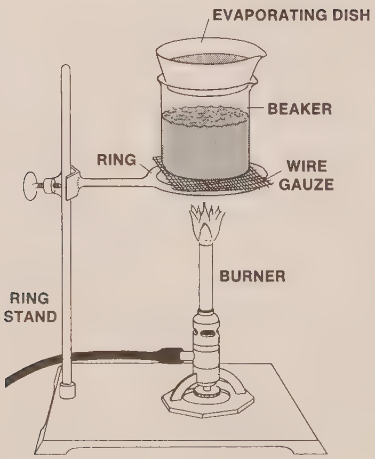
FOLDED FILTER



FILTER IN FUNNEL

Set up apparatus as shown. Place liquid to be evaporated in evaporating dish.





DATA TABLE

PROPERTIES OF SOME ELEMENTS

Name	Phase At Room Temperature	Density (g/cm³)	Density (g/l)	Melting Pt. (°C)	Boiling Pt. (°C)	Appearance
Aluminum	solid	2.7		660.2	2467	silvery
Bismuth	solid	9.8		271.3	1560	silvery-white or reddish
Bromine	liquid	3.1		−7.2	58.8	dark red
Chlorine	gas		2.5	−103	−34.6	greenish-yellow
Gold	solid	19.3		1063	2600	yellow
Helium	gas		0.18	−272 (at high pressure)	−268.9	colorless
Hydrogen	gas		0.09	−259	−268.9	colorless
Iodine	solid	4.93		113-114 sublimes	184.4	silvery-blackish solid violet gas
Lead	solid	11.3		327	1515	silvery-black
Magnesium	solid	1.74		651	1107	silvery-white
Oxygen	gas		1.1	−218.4	−183.	colorless
Sulfur	solid	2.0		114.5	444.6	yellow
Zinc	solid	7.14		419.5	907	bluish-white, lusterous

sublimes: turns directly from solid to gas

Helium's melting point is determined under 26 atm of pressure

(Data from *Chemistry and Physics Handbook*, 46th ed.)

Laboratory Experiments in

ACTION CHEMISTRY

UNIT II

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Well, change your partners and doe-zee-doe! Make two new couples from two of the old.

Chemical "Drop-Outs"

Many chemical reactions take place in solution. Solid chemicals react more quickly when dissolved than they do as solids. In this experiment you may see some dissolved chemicals react when solutions are combined. New substances may be produced which have very low solubility. Substances having low solubility drop out of solution. In a chemical reaction:

substances which react to form new substances are called REACTANTS;
new substances produced are called PRODUCTS;
substances which fall out of solution as solids when they are products of a reaction, are called PRECIPITATES.

Experiment 5.1

Equipment and Supplies per Lab Station

(6) small test tubes; test tube rack
1 set of five solutions in 30 ml dropping bottles:

silver nitrate
barium chloride
iron (III) chloride
potassium sulfate
sodium carbonate

What to do

1. Obtain solutions of the following chemicals, each contained in a separate dropping bottle.
   silver nitrate
   barium chloride
   iron (III) chloride
   sodium carbonate
   potassium sulfate
2. Put 10 drops of silver nitrate solution in a small test tube numbered "1". Add 10 drops of barium chloride solution. Record your observations in the Data Chart.
3. Number nine additional small test tubes. Make all of the combinations listed in the Data Chart. Use 10 drops of each solution as you did in step 2. Record your observations in the Data Chart.
4. Write the word equation in the second chart for each combination in which you see a precipitate form. Fig. 19 will help you work out the products of the reaction.

5. Look up the name of each possible new substance in the Solubility Reference Table. Place an arrow, ↓, in the equation after the name of the product of low solubility. The arrow shows that the compound has dropped out as a precipitate.

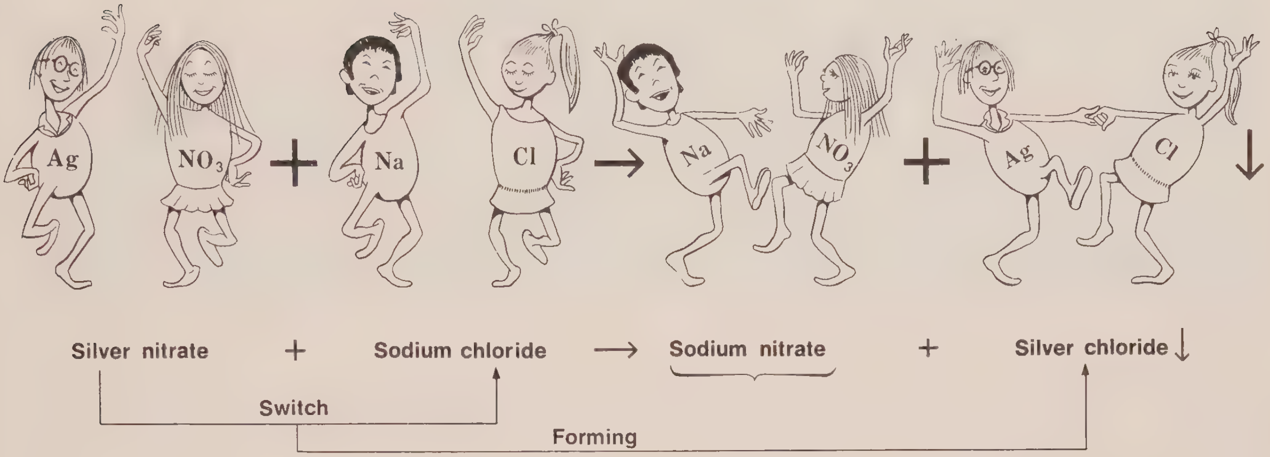

Ag
NO3
Na
Cl
Na
NO3
Ag
Cl
Silver nitrate + Sodium chloride → Sodium nitrate + Silver chloride↓
Switch
Forming

Fig. 19. A double replacement reaction.

DATA CHART

Pair No.	Solution I	Solution II	Observations
1.	silver nitrate	barium chloride	
2.	silver nitrate	sodium carbonate	
3.	silver nitrate	iron (III) chloride	
4.	silver nitrate	potassium sulfate	
5.	barium chloride	sodium carbonate	
6.	barium chloride	iron (III) chloride	
7.	barium chloride	potassium sulfate	
8.	sodium carbonate	potassium sulfate	
9.	iron (III) chloride	potassium sulfate	
10.	iron (III) chloride	sodium carbonate	

SOLUBILITY REFERENCE TABLE

	carbonate	chloride	nitrate	sulfate
barium	i	s	s	i
iron (III)	n	s	s	Ss
potassium	s	s	s	s
silver	i	i	s	Ss
sodium	s	s	s	s

i-very low solubility (nearly insoluble)
Ss-slightly soluble
s-soluble
n-not isolated (iron (III) carbonate has not been found to exist)

Questions to answer

1. Suppose you have unmarked bottles of water, sodium carbonate, silver nitrate, and barium chloride, could you tell which was which by looking at the bottles?

   Explain. ____

   ____

2. How could you tell the difference between unmarked bottles of barium chloride and silver nitrate? ____________

   ____

   ____

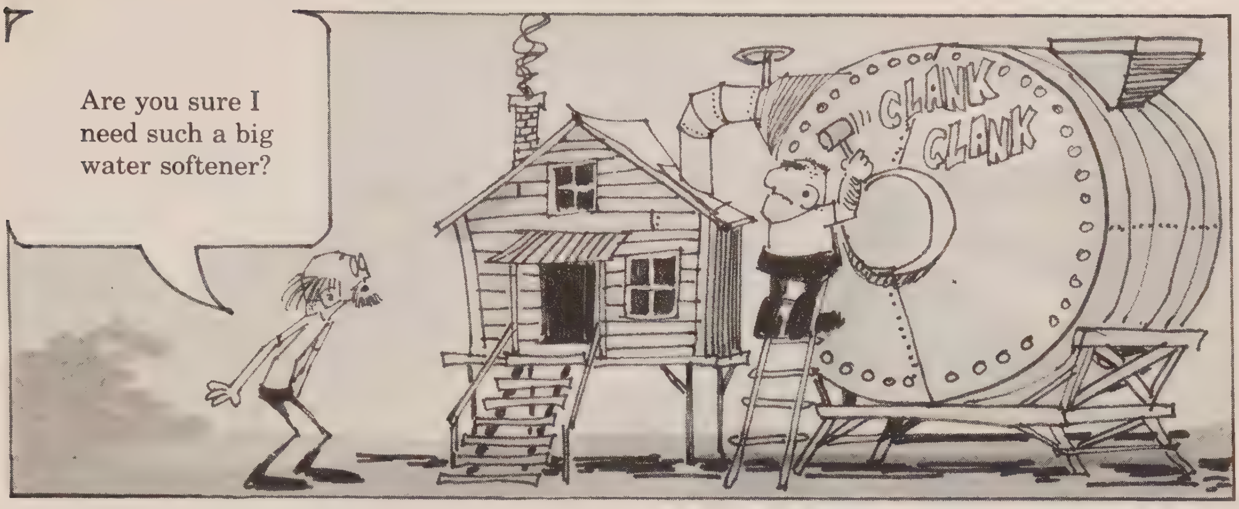
3. Look at the Solubility Reference Table. Which ones of the possible reactants always form soluble combinations? ______

   ____

   ____

EQUATION CHART

Pair No.	Word Equation for the Reaction
1.	
2.	
3.	
4.	
5.	
6.	
7.	
8.	
9.	
10.	


Are you sure I need such a big water softener?
CLANK
CLANK

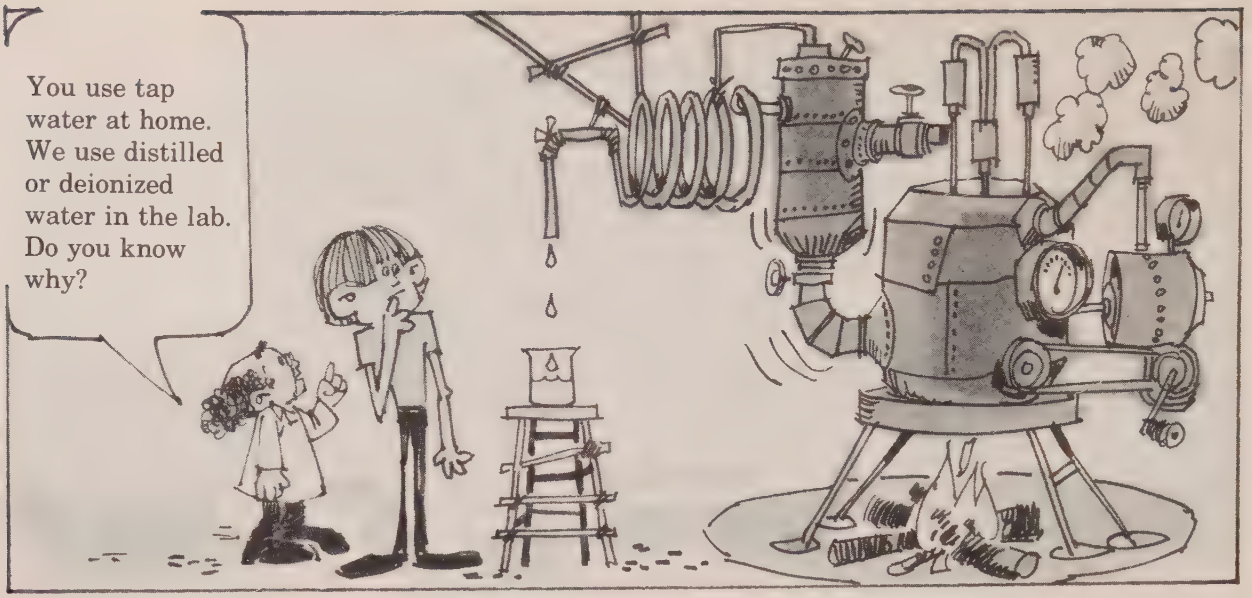
An Exercise in Writing Double Exchange Word Equations

Directions

1. Look up the solubility of each compound given below in the "Table of Solubilities in Water".
2. Leave the right hand side of the equation blank if a combination of the two compounds would NOT produce a precipitate.
3. Complete the word equation if a precipitate should form.
4. Look up the solubility of the compounds produced.
5. Place an arrow beside the name of the compound which is the precipitate, as you did in Experiment 5.1.

EQUATIONS

1. sodium bromide + barium chloride → ______________ + ______________
2. calcium nitrate + sodium carbonate → ______________ + ______________
3. ammonium chloride + sodium sulfate → ______________ + ______________
4. potassium iodide + lead nitrate → ______________ + ______________
5. magnesium iodide + sodium hydroxide → ______________ + ______________
6. aluminum sulfate + potassium hydroxide → ______________ + ______________
7. copper chloride + zinc nitrate → ______________ + ______________
8. lead nitrate + sodium sulfate → ______________ + ______________
9. silver nitrate + sodium chloride → ______________ + ______________
10. zinc chloride + potassium carbonate → ______________ + ______________


You use tap water at home. We use distilled or deionized water in the lab. Do you know why?

Chemists are Choosy About Water

You use tap water at home, but a chemist uses distilled water or deionized water in the laboratory.

A WATER SOLUTION is water containing dissolved chemicals. When you dissolve table salt in water, you are making a solution of salt in water.

Why does a chemist use distilled or deionized water to make water solutions? In this experiment you may discover the reason.

Experiment 5.2

Equipment and Supplies per Lab Station

(4) test tubes, 18 × 150 mm; plastic teaspoon
(4) watch glasses; beaker—250 ml
ring stand with ring and wire gauze; burner
20 ml distilled water; 20 ml deionized water
20 ml sodium chloride solution
1 g, approx., of lead nitrate crystals

What to do

PART I

1. Number four test tubes, 1, 2, 3, and 4 respectively. Pour 10 ml of each of the liquids listed in the Data Chart into one of the four separate test tubes.

DATA CHART

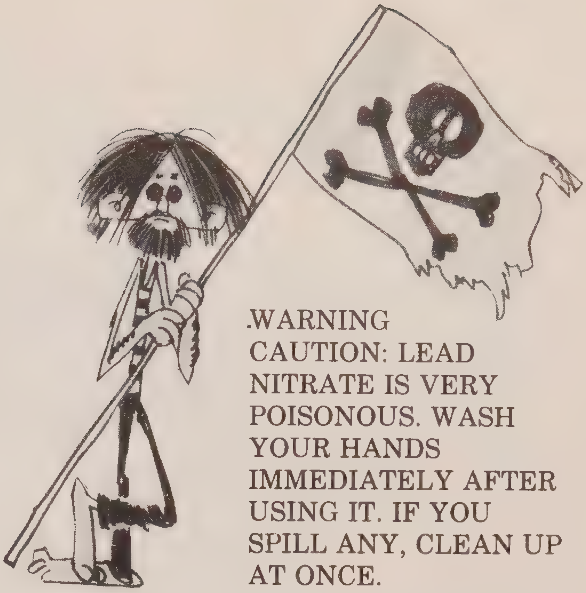
Sample	Observations after adding lead nitrate crystals
Test tube 1—tap water	
Test tube 2—distilled water	
Test tube 3—deionized water	
Test tube 4—solution of table salt	

2. To each test tube add one-fourth teaspoonful of lead nitrate crystals.
3. (a) Compare the results in the four tests using the Data Chart on page 6. ______

______

(b) What do you think could be causing any difference in appearance in the four tubes? ______

______



WARNING
CAUTION: LEAD NITRATE IS VERY POISONOUS. WASH YOUR HANDS IMMEDIATELY AFTER USING IT. IF YOU SPILL ANY, CLEAN UP AT ONCE.

PART II

1. Set up an evaporating apparatus as shown in Fig. 20.

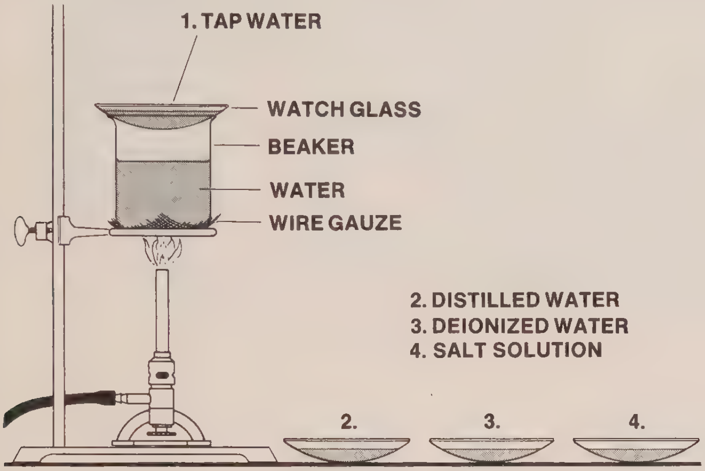


Fig. 20. Set up for Experiment 5.2.

2. Evaporate 5 ml of each type of water and of the salt solution. What was left on the watch crystal after each was dry?

From the tap water? ______

From the distilled water? ______

From the deionized water? ______

From the salt solution? ______

3. Do the results of these tests help to explain any difference in appearance of the tests in PART I? ______

Questions to answer

1. From your results in PART II, what explanation do you see for any difference in appearance in the various solutions of lead nitrate in PART I? ______

______

2. Do you have any other explanation of possible differences?

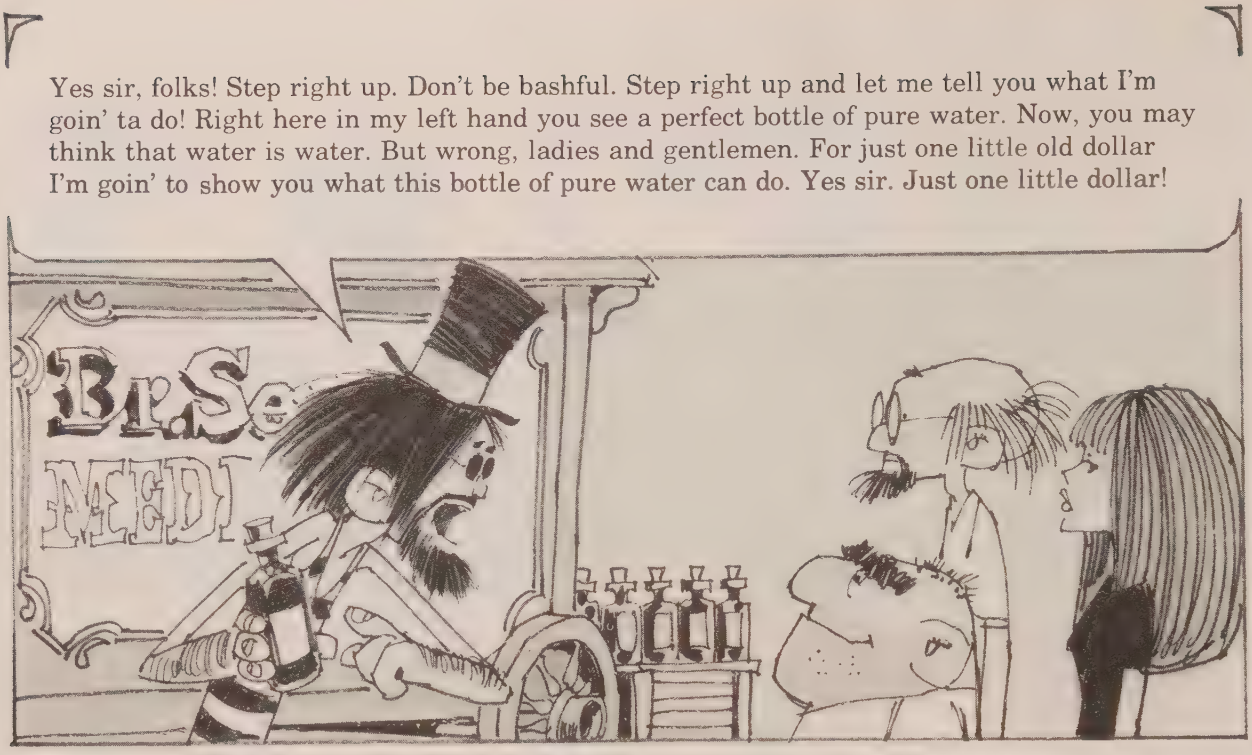
______

3. Why do you think chemists use distilled or deionized water in the laboratory solutions? ______

Extra for experts

Why is distilled water used to refill automobile batteries? ______

______


Yes sir, folks! Step right up. Don't be bashful. Step right up and let me tell you what I'm goin' ta do! Right here in my left hand you see a perfect bottle of pure water. Now, you may think that water is water. But wrong, ladies and gentlemen. For just one little old dollar I'm goin' to show you what this bottle of pure water can do. Yes sir. Just one little dollar!
Dr. Se
MEDI

THE RELATIVE HARDNESS OF WATER SAMPLES

Some dissolved minerals cause water to form precipitates, curds with soap, and a rock-like deposit in boilers. In this experiment you will compare the amount of dissolved minerals, or hardness, in several samples of water.

Experiment 5.3

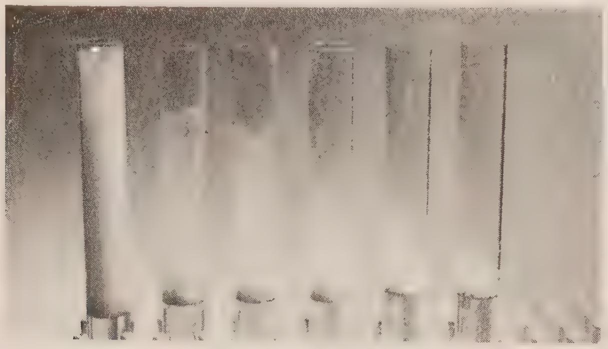
Equipment and Supplies per Lab Station

graduated cylinder–10 ml or 25 ml
(6) test tubes; stopper to fit the test tubes
test tube rack; glass marking pen or pencil
distilled water; samples of hard water
water from as many different sources as possible
standard soap solution—1 ml for each sample tested

What to do

1. Pour 10 ml of distilled water into a test tube. This will be the control test tube. Mark it #1.
2. Add 1 ml of standard soap solution.
3. Stopper the test tube. Shake vigorously for two minutes. Set the test tube in the rack.
4. Mark the height of the suds with a glass marking pen or pencil.

5. Pour 10 ml of tap water into a second test tube. Mark it #2. Repeat steps 2, 3, and 4.
6. Use 10 ml of other samples such as rain water, river water, pond water, or samples of hard water made in the laboratory. Repeat steps 2, 3, and 4. Number the test tubes #3, #4, etc.
7. Arrange the test tubes in order of suds height.



DATA CHART

	No. of tube	Type of water tested
(a) Most suds		
(b)		
(c)		
(d)		
(e) Least suds		

8. Write down any other observations you make. ____________

____

____

Questions to answer

1. Which sample do you believe contained the most dissolved minerals? ____________

   What made you think this? ____________

   ____

2. Why do you think that the pharmacist uses distilled water when water is required in making up a prescription? ____________

   ____

3. Why is distilled water recommended for use in most steam irons? ____________
4. Why do some women like to use rain water or melted snow to shampoo their hair? ____________
5. What makes the ring around the bath tub? ____________

   ____

Hey, let's go check our distillery!

PURIFYING WATER FOR THE CHEMISTRY LABORATORY

Two methods are in general use for purifying water for laboratory work.

1. *Distillation,* in which two phase changes you have already studied take place.
    - (a) *Evaporation:* A liquid is changed rapidly to a vapor (gas) by direct heating.
    - (b) *Condensation:* The vapor is cooled in a *condenser,* such as the one in Fig. 21, and changed back to a liquid in a separate container.
    The product of any distillation is called the DISTILLATE. The distillate of water is DISTILLED WATER.
2. *Deionization,* removes mineral substances dissolved in a liquid as the liquid passes down through a column of a special chemical. Water that has passed through such a column is called DEIONIZED WATER.

In this experiment you will make distilled water and also deionized water.

Experiment 5.4

Equipment and Supplies per Lab Station

beaker—150 ml; (4) flasks, Erlenmeyer—125 ml or 250 ml
distilling flask—250 ml,.with a one-hole stopper for thermometer
thermometer; condenser, with stopper bored to fit side arm of distilling flask
graduated cylinder—50 ml or 100 ml; deionizer column (see note below)
(2) stoppers for two of the Erlenmeyer flasks
ring stand with clamp and ring with wire gauze for the distilling flask
ring stand with a large clamp for the condenser
ring stand with a large clamp for the deionizer column
  Note: It is not practical to attach the deionizer column to either of the ring stands being used for the distillation apparatus.
burner; balance
2 grams of one of a number of different colored salts
distilled water

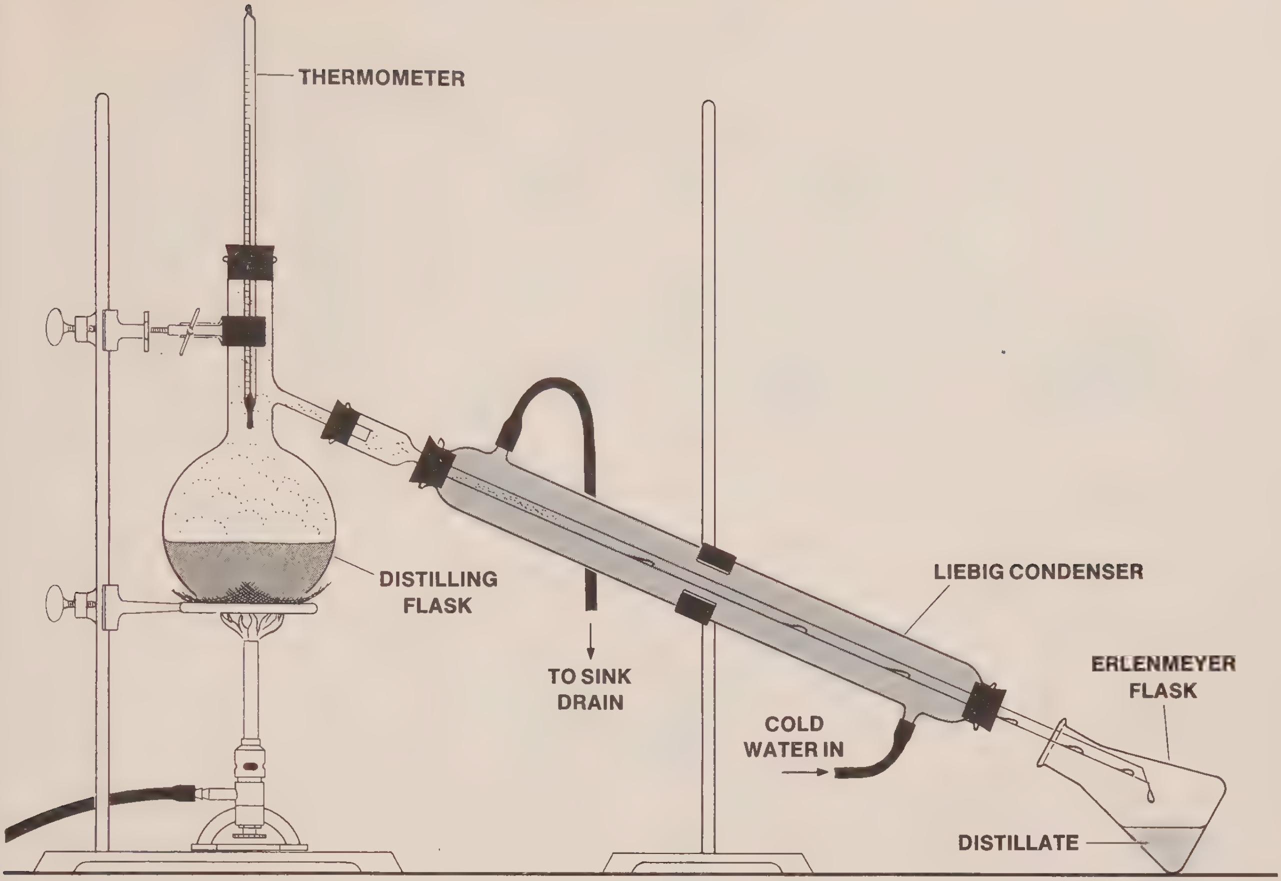


Fig. 21. Distillation set up.

What to do

PART I. Distillation

1. Dissolve about 2 g of the colored chemical assigned to you by your teacher in 200 ml of distilled water. The colored chemical is to be the impurity in the water. See Fig. 22.

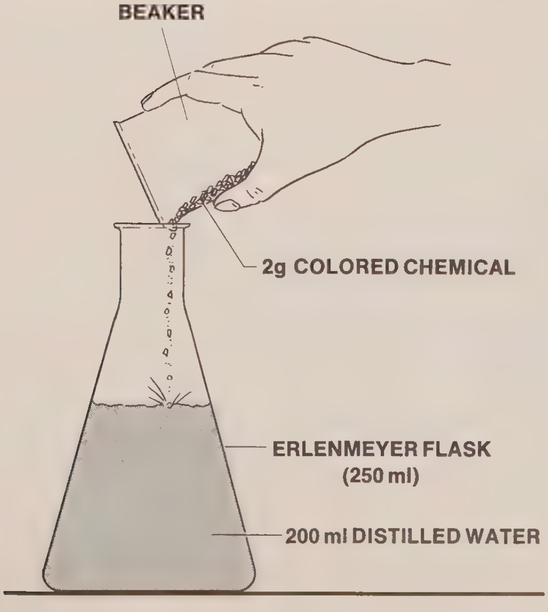


Fig. 22. Making the sample to be purified.

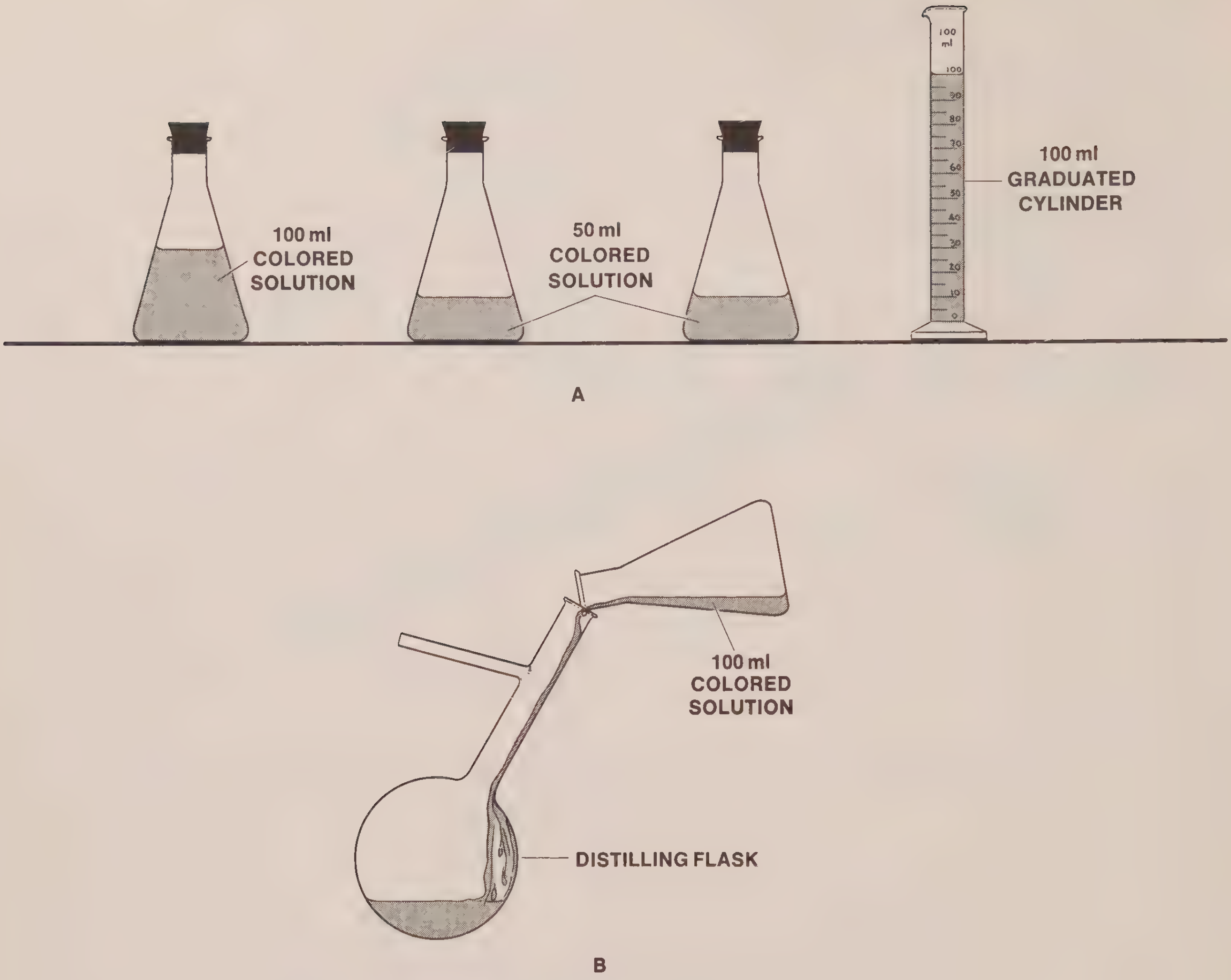


Fig. 23.

2. Divide the colored water into *one* 100 ml portion and *two* 50 ml portions. Put the two 50 ml portions in separate flasks. Stopper the flasks and set them aside for PART II. See Fig. 23.
3. Assemble the apparatus as shown in Fig. 21. Pour the 100 ml portion of the colored water into the distilling flask. Replace the stopper.
4. Ask your teacher to approve the apparatus. Turn on the water to circulate through the condenser.
5. Heat the flask. When boiling begins, adjust the flame so that there is gentle, steady boiling in the flask.
6. Record the temperature as boiling begins. ____________°C
7. Collect about 50 ml of the distillate, which is distilled water.
8. Turn off the burner flame. Turn off the water circulating in the condenser.

PART II. Deionization

1. Set up a deionizer column as shown in Fig. 24.
2. Pour one of the 50 ml portions of colored water set aside in PART I, into the plastic cup at the top of the deionizer column. Collect the water as it drips from the bottom of the column.
3. Pour the water collected through the deionizer again. Repeat this process several times. Save the final amount collected.

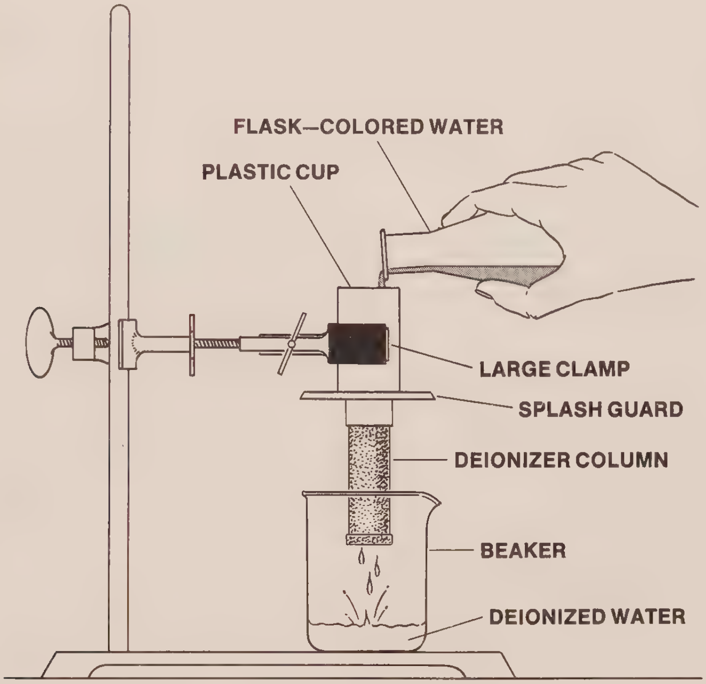


Fig. 24. Deionizer column.

Questions to answer

PART I. Distillation

1. How does the color of the distillate compare with the color of the 50 ml samples set aside for PART II? ______________
____
2. How can you account for any difference in color? _________
____
3. What two phase changes occur during distillation?
(a) ____________________ (b) ____________________
4. How would you define distillation in your own words? ____
____
____
____

Extra for Experts

Try to explain the temperature at which boiling in the distilling flask occurs. You may need to look this up in reference books. ____________________

____

____

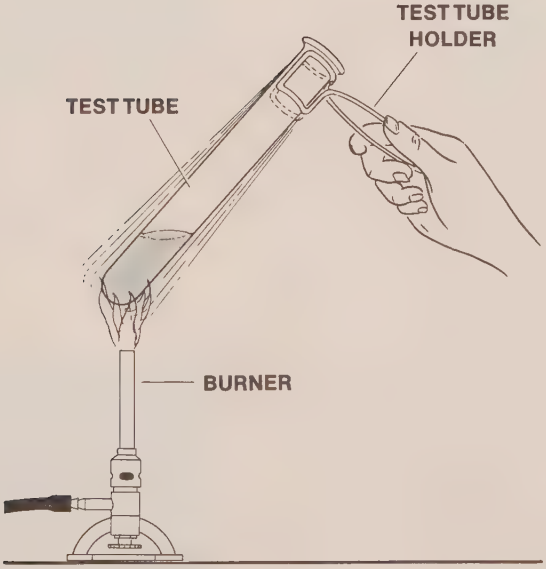
PART II. Deionization

1. How does the water collected from the deionizer column compare in color with
   (a) the second 50 ml sample from PART I? ____________________
   ____
   (b) the distilled water from PART I? ____________________
   ____
2. What do you think happens to the colored chemical as the solution goes through the deionizer column? ____________________
   ____
   ____
3. Other students in the class may have used different colored chemicals. If so, was there any difference that you could see in either the distilled water or the deionized water from different colored solutions? ____________ If so, explain. ____________________
   ____

HOW TO

Heat Material in a Test Tube

1. Always use a test tube holder or clamp when heating a test tube.
2. Never point a heated tube at anyone because the liquid may splash out of the test tube.
3. Do not heat any one spot on the test tube continuously. Move the test tube as shown.



Experiment 6.1

Equipment and Supplies per Lab Station

(4) beakers—100 ml; graduated cylinder—25 ml
(6) test tubes; (2) stoppers to fit the test tubes
test tube rack; mortar and pestle; laboratory burner
test tube holder
(6) large crystals of a colored salt such as
copper sulfate or nickel sulfate

A Chemical Dissolving Race

Sometimes solid substances dissolve quickly. At other times the dissolving is slow. Watch how fast the dissolving occurs in this chemical race. It may help you find ways to save time when you make solutions.

ON YOUR MARK or getting the contestants ready:

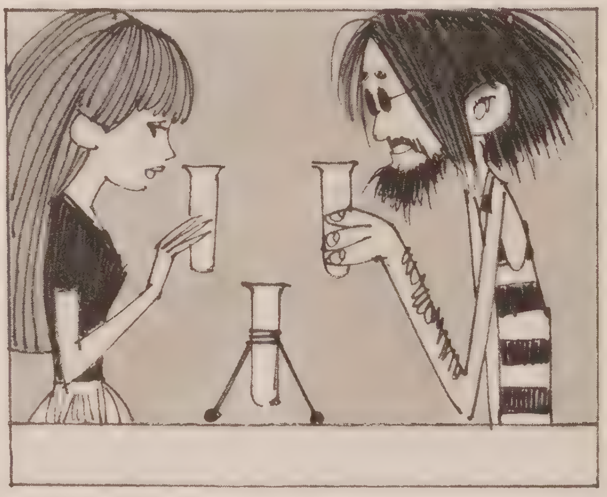
1. Select four large crystals of copper sulfate or nickel sulfate which are about the same size and which can just go into a test tube. Label four test tubes, 1, 2, 3, and 4.
2. Place one of the crystals in test tube numbered 1. Put another crystal in test tube numbered 3.
3. Pulverize the other crystals in a mortar. Place half of the powdered solid in test tube 2 and the other half in test tube 4.

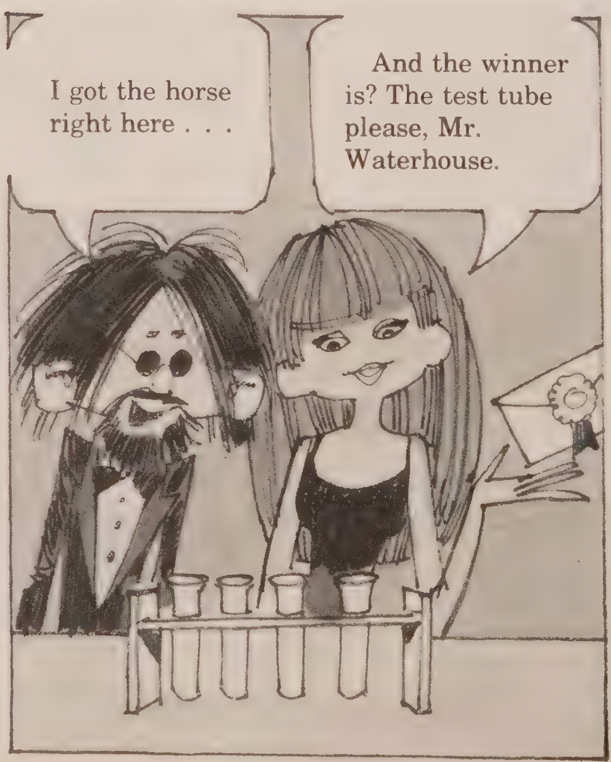
GET SET or have the race course ready for the contest:

4. Place 5 ml of water in each of four small beakers.
5. Have two stoppers for test tubes 3 and 4 ready to use.
6. One partner should hold two beakers containing the water samples ready to pour into test tubes 1 and 2. The other partner should be ready to pour the water samples into test tubes 3 and 4.

GO! or give the starting signal:

7. When you are both ready, give the starting signal and each partner should pour the water into the test tubes.




I got the horse right here . . .
And the winner is? The test tube please, Mr. Waterhouse.

8. Let test tubes 1 and 2 remain undisturbed. Stopper test tubes 3 and 4 and shake them steadily for one minute.

WATCHING THE RACE or making observations:

9. Observe the samples so you can declare which sample dissolves first. If not all dissolves, the depth of color can be a clue that shows how much has dissolved.
10. Declare the winner or winners by listing the order in which the contestants crossed the "finish line."

Winner ____________ Third place ________

Second place ____________ Fourth place ________

11. Refer to the "How To" page on heating a liquid in a test tube safely.
12. Label two test tubes 5 and 6. Choose two colored crystals of the same size. Place one in each test tube. Add 5 ml of water to each of the two test tubes at the same time. One partner shakes test tube 5 vigorously. The other partner heats test tube 6 gently. In which test tube did the crystal dissolve first? ____________

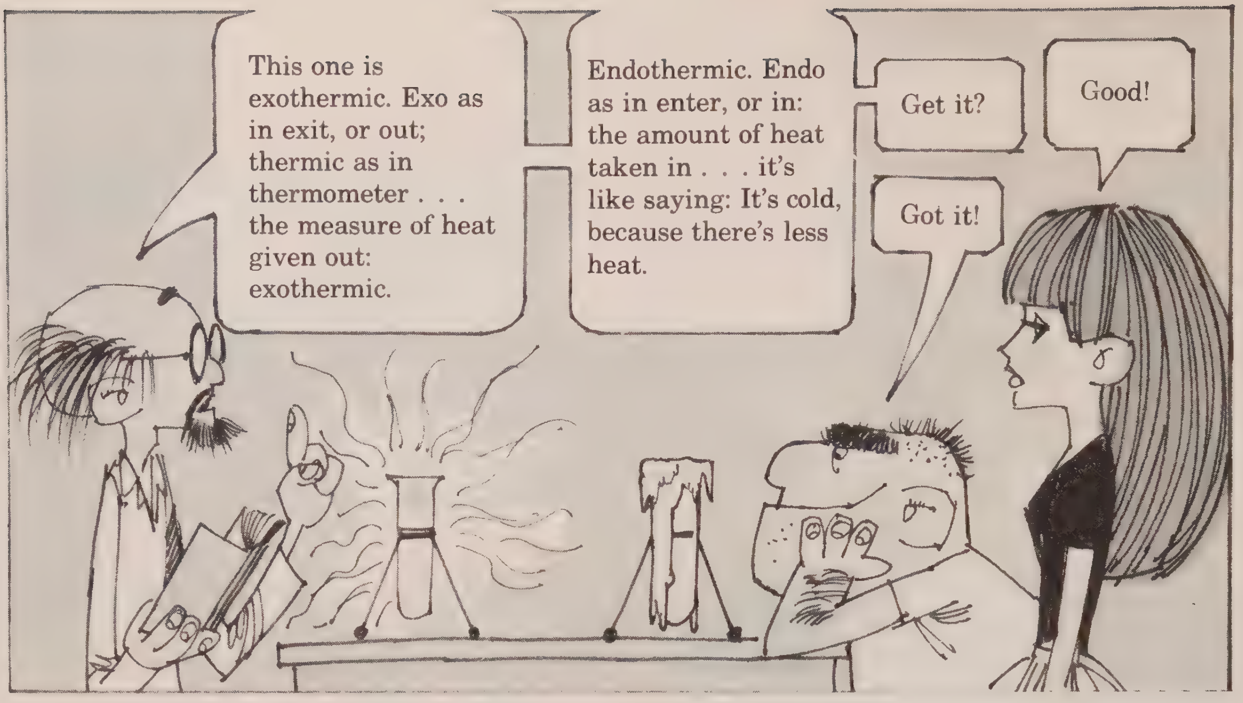
EVALUATING THE RACING RECORD or making some hypotheses:

13. How can you account for the differences in the rate of dissolving
    - (a) When the crystal was powdered? ____________

      ____
    - (b) When the large crystal was shaken steadily? ________

      ____
14. What did you observe about the effect of heat on the rate of dissolving? ____________

    ____


This one is exothermic. Exo as in exit, or out; thermic as in thermometer . . . the measure of heat given out: exothermic.
Endothermic. Endo as in enter, or in: the amount of heat taken in . . . it's like saying: It's cold, because there's less heat.
Get it?
Good!
Got it!

HEAT AND DISSOLVING

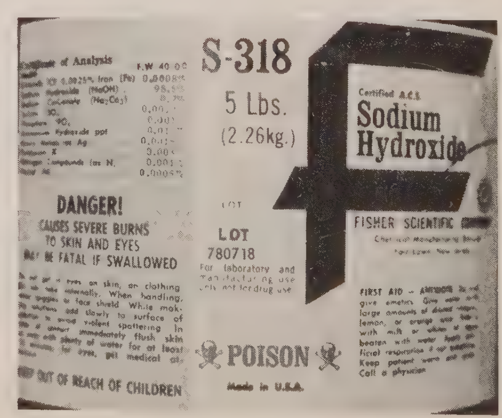
Two things happen when a substance dissolves in water. Heat energy is being absorbed for use in one change and at the same time heat energy is being released in a second change. If the amount of heat taken in and the amount given out are about equal, there will be little or no change in temperature during the dissolving. The solution becomes hot if the amount of heat released is greater than the amount absorbed. The solution becomes cold if the amount of heat absorbed is the greater quantity.

EXOTHERMIC DISSOLVING: the solution becomes hot.

ENDOTHERMIC DISSOLVING: the solution becomes cold.

All of this may sound a little complicated, but it is really simple to observe as you will see in this experiment.

Experiment 6.1 — A Extension



Equipment and Supplies per Lab Station
(3) test tubes; graduated cylinder; thermometer
test tube rack; forceps; teaspoon-size scoop
distilled water; 3 pellets of sodium hydroxide
5 grams of ammonium nitrate; 5 grams of potassium nitrate

What to do
1. Pour 10 ml of distilled water into each of three test tubes.
2. Take the temperature of the water in each test tube. Record the readings in the Data Chart.

3. Read the label on the bottle of sodium hydroxide in the illustration.
   (a) Why must you use forceps to pick up sodium hydroxide pellets? ______________________
   (b) What other precautions must you observe when working with sodium hydroxide pellets or solution? ____
   ____
   ____
4. Add three pellets of sodium hydroxide to test tube 1 with *a forceps.* Place a thermometer in the test tube. Stir the pellets and water slowly and gently. Watch the thermometer closely. Record in the Data Chart the temperature at the point of greatest change. Record any other observations below.
5. Repeat step 4. Use test tube 2. Add 5 g (one spoonful) of potassium nitrate in place of the pellets. Stir gently. Again record the temperature at point of greatest change. Are there any other observations?
6. Add 5 g (one spoonful) of ammonium nitrate to test tube 3. Stir gently. Once more record the temperature at the point of greatest change. Are there any other observations?

DATA CHART

Test Tube	Temperature of Water	Temperature at Point of Greatest Change	Increase or Decrease?	Exothermic or Endothermic?
#1				
#2				
#3				

7. Other observations
   Test tube 1 ______________________
   Test tube 2 ______________________
   Test tube 3 ______________________

Questions to answer

Write in your own words what you have found out about heat and dissolving. ______________________

____

____

Extra for experts

What new experiment about dissolving would you like to try? ______________________

____

____

These bottles of soda carry dissolved gas. Gases can dissolve in liquid.
Oh, yeah! I thought only hard things like sugar and salt dissolve in liquids! How do you like that. The bottle is full of gas!
PEPSI

Solubility of Gases

What are the conditions affecting the solubility of carbon dioxide in soda pop?

Would the solubility of other gases be affected in the same way by the same conditions?

Would the solubility of gases be changed if the conditions were changed?

See what you can find out about these questions from a bottle of soda pop!

Demonstration 6-A

Equipment and Supplies per Demonstration

(2) ice-cold bottles of soda water; (2) beakers—400 ml burner; ring stand with ring and wire gauze

What to do

1. Inspect two cold bottles of soda water. Do you see any evidence that carbon dioxide is there? ________________
2. Uncap one bottle and write down any observation. ________________
3. Put a pressure cap back on the bottle. What happens to any action in the bottle? ________________
   How can you account for this? ________________
4. Open the other bottle and pour half of the soda water into one beaker, the other half into another beaker. Compare the amount of effervescence in the two beakers. ________________

5. Put one of the beakers on an iron ring on a ring stand. Heat the soda water gently.

   Compare the effervescence in the two beakers. ____________

   ____

   How can you account for any differences? ________________

   ____

6. This experiment shows the following facts about the solubility of gases.
   (a) As the pressure on the solution of the gas increases, the solubility of the gas ________________________.
   (b) As the temperature of the solvent increases, the solubility of the gas ________________________.

Questions to answer

1. Sometimes a bottle of pop standing in the sunshine will explode. What causes the explosion? ________________
2. Why does an opened bottle of ginger ale go "dead" after a period of time? ________________________________
3. Why do you think that trout die when they are taken from a running brook and kept in a balanced aquarium where the water temperature is higher? ________________

   ____

   ____

4. What change in the amount of dissolved oxygen present occurs as water is sprayed into the air or as it passes over rapids? ____

   ____

   ____

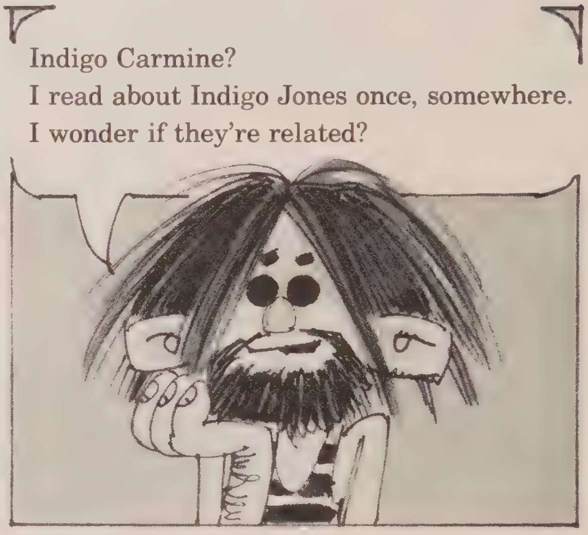
   ____

   ____

Indicator Dye Solutions

Water is used more than any other liquid to make laboratory solutions. However, certain other liquids may be used as solvents.

In later experiments you will need solutions of the indicator dyes, phenolphthalein and indigo carmine. Can these dyes be dissolved in water? If not, what liquid can be used?


Indigo Carmine?
I read about Indigo Jones once, somewhere.
I wonder if they're related?

Demonstration 6-B

Equipment and Supplies per Demonstration

(4) flasks—250 ml; graduated cylinder—100 ml
(4) stoppers to fit the flasks
2 g of phenolphthalein; 0.5 g of indigo carmine dye
200 ml of distilled water; 200 ml of 50% ethanol
10 ml of dilute hydrochloric acid solution
10 ml of dilute sodium hydroxide solution

What to do

1. Weigh two 1 g samples of phenophthalein. Place one of the weighed samples in each of two 250-ml flasks.
2. Add 100 ml of distilled water to one of the flasks. Stopper the flask and shake it vigorously. Let it stand for one minute.

   What has happened? ______________________________
3. Add 100 ml of 50% ethanol solution to the second flask. Stopper this flask, shake it, and let it stand. Is the result the same with the ethanol as with the distilled water? ______________ If not, how is it different? ______________________________

4. Weigh two 0.25 g samples of indigo carmine. Place one of the samples in each of two 250-ml flasks.
5. Repeat steps 3 and 4.
   (a) What happens with the indigo carmine and distilled water? ____________________
   (b) What happens with the indigo carmine and the 50% ethanol? ____________________
6. What do you conclude is the proper solvent for phenolphthalein and for indigo carmine? ____________________

   ____

   ____

7. Number four test tubes 1, 2, 3, and 4.
8. Place in both test tube 1 and 2, 5 ml of dilute acid solution.
   (a) To test tube 1 add two drops of the phenolphthalein solution. What happens? ____________________

   ____

   ____

   (b) To test tube 2 add two drops of the indigo carmine solution. What happens? ____________________

   ____

   ____

9. In test tubes 3 and 4 each, place about 5 ml of dilute sodium hydroxide solution.
   (a) To test tube 3 add two drops of the phenolphthalein solution. What happens? ____________________

   ____

   ____

   (b) To test tube 4 add two drops of the indigo carmine solution. What happens? ____________________

   ____

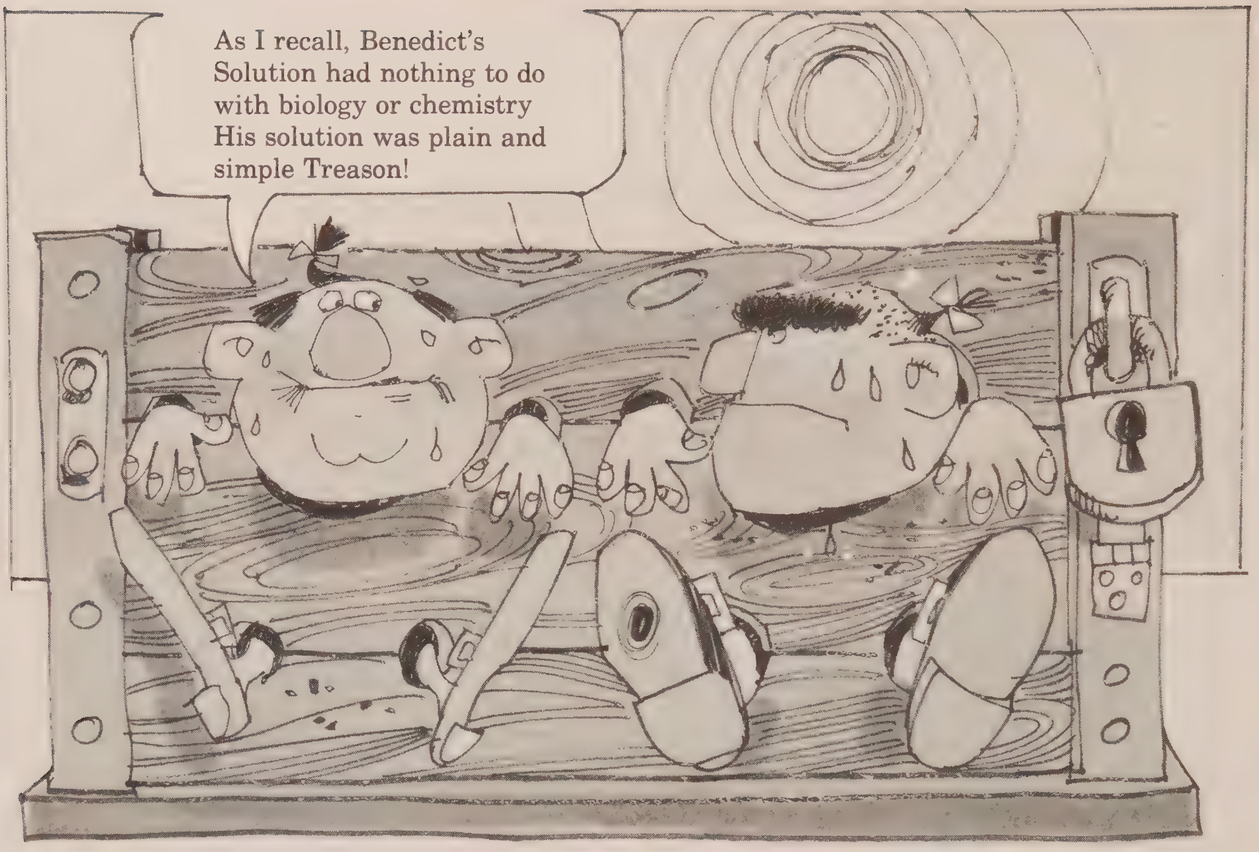
   ____

10. How do you think these dye solutions might be used in the chemistry laboratory? ____________________

    ____

    ____

11. Transfer the phenolphthalein solution and the indigo carmine solution to properly labeled small dropper bottles. Store in the stockroom for future use.


As I recall, Benedict's Solution had nothing to do with biology or chemistry His solution was plain and simple Treason!

Benedict's Solution

Benedict's Solution is used to test for the presence of simple sugars. You will use it in later work.

In Benedict's Solution three different chemicals, instead of only one, are dissolved in the same solution. The directions for combining these chemicals must be followed with great care.

Experiment 6.2

Equipment and Supplies per Lab Station

beaker—250 ml; beaker—400 ml; graduated cylinder—100 ml
flask—250 ml (with stopper); glass stirring rod
medicine dropper; test tube; balance; burner
ring stand with ring and asbestos wire gauze
test tube holder or clamp
distilled water; 1.7 g of hydrated copper sulfate
10.0 g of sodium carbonate; 17.3 g of sodium citrate

What to do

1. *Solution 1*
   - (a) Weigh on separate waxed weighing papers
     17.3 g of sodium citrate
     10.0 g of sodium carbonate
   - (b) Place weighed chemicals together in one 400-ml beaker. To this add 80 ml of distilled water. Warm the solution gently until the chemicals dissolve. Cool for a few minutes.
   - (c) Pour the solution into a 100-ml graduated cylinder. Add enough water to make 85 ml. Pour the solution back into the beaker. This is Solution 1. Set it aside in a safe place.

2. *Solution 2*
   (a) Weigh 1.7 g of copper sulfate on a waxed weighing paper.
   (b) Place this in a second beaker. Add 10 ml of distilled water. Stir to help dissolve the copper sulfate. If necessary, the solution may be warmed slightly to speed up dissolving.
3. Pour the copper sulfate solution *slowly, with constant stirring* into Solution 1. See Fig. 25. Transfer the combined solutions to the 100-ml graduated cylinder. Add enough distilled water to make 100 ml of solution.

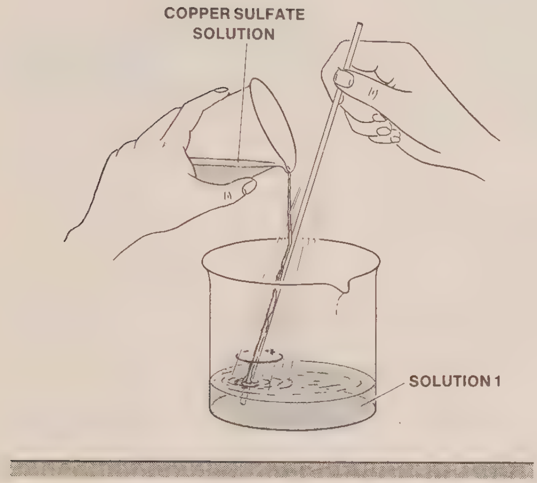


Fig. 25.

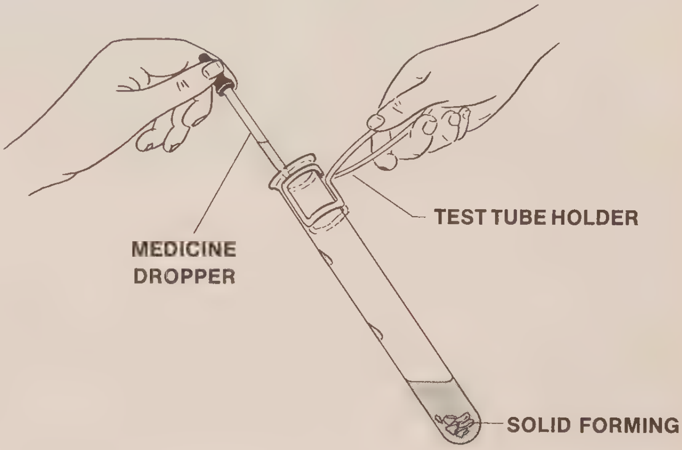


Fig. 26. Testing sugar solution with Benedict's solution.

4. Pour the solution into a 250-ml flask. Stopper the flask. Shake well to mix the solution. The solution should be deep blue and perfectly clear.
5. Testing your solution.
   (a) Place 5 ml of the solution in a test tube. Heat just to boiling.
   (b) Add one to five drops, one drop at a time, of corn syrup. See Fig. 26. If a brick red solid forms, you have made the Benedict's solution correctly.

   Note: The forming of the brick red solid with Benedict's solution is a positive test for the presence of a simple sugar. The sweetness of corn syrup is due to dextrose, a simple sugar.
6. Turn your Benedict's Solution in to your teacher if it gives a satisfactory test. It will be stored for later use.

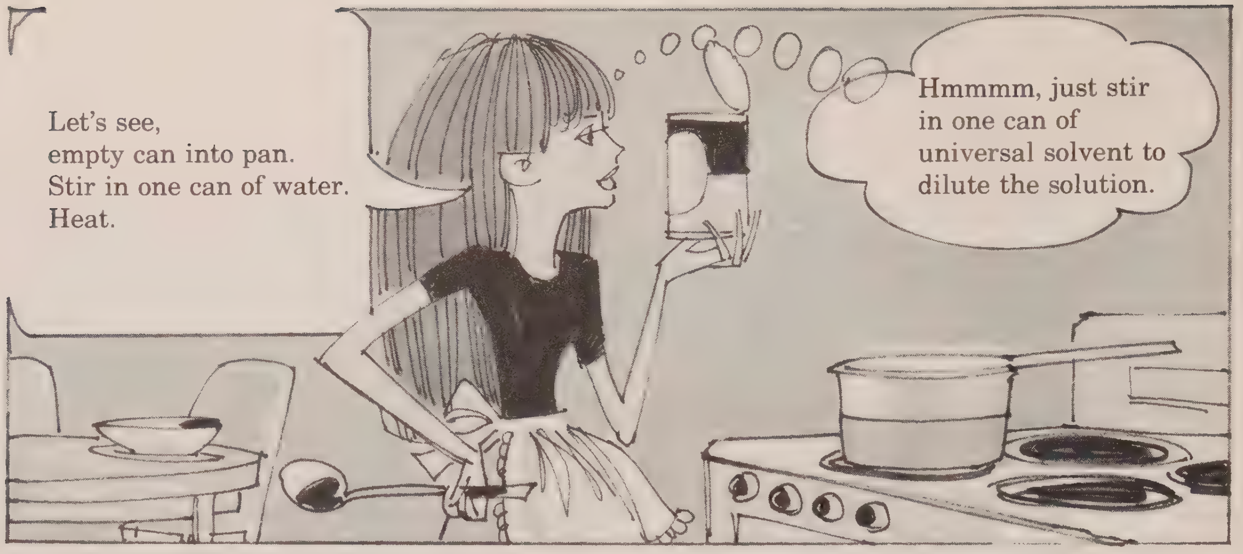
Questions to answer

1. What changes did you notice as the copper sulfate solution was added to the Solution 1? ____________

   ____
2. How would you test unsweetened fruit juice for simple sugar? ____________

   ____

   ____



Make it PALER, Make it **DARKER**

In this experiment you will change the amount of dissolved solute per milliliter of solvent to make a solution either more dilute or more concentrated.

Experiment 6.3

Equipment and Supplies per Lab Station

beaker—150 ml; evaporating dish; graduated cylinder—25 ml
(3) test tubes; burner; ring stand with ring and asbestos wire gauze
40 ml of the solution of a colored salt; distilled water

What to do

1. Obtain 40 ml of a colored solution provided by your teacher.
2. Pour 10 ml of the colored solution into a test tube. Set the test tube aside. It will be used as a standard for comparison.
3. Mix 10 ml of water with 10 ml of the colored solution in a small beaker.
4. Pour 10 ml of the colored solution you just made into a test tube.
5. Compare the intensity of the colors of the standard solution and of the one just made.
   (a) Describe what you see. ____________________
   ____
   (b) Which one has the greater amount of solute per milliliter of solution? ____________________
   (c) What makes you think this? ____________________
   (d) How can you make any solution more dilute (less solute per milliliter of solution)? ____________________

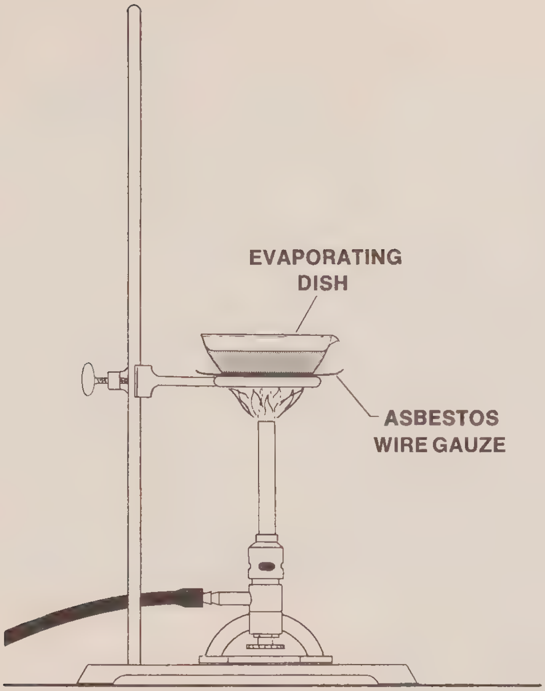


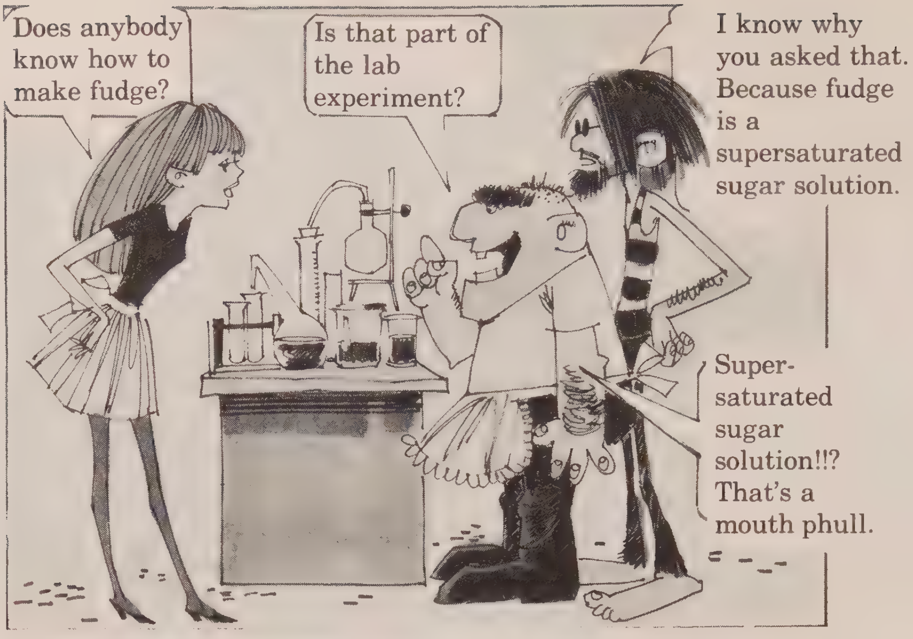
Fig. 27.

6. Pour 20 ml of the standard colored solution into an evaporating dish.
7. Heat the solution slowly (Fig. 27.) until there is about 10 ml of solution remaining in the dish.

   What leaves the dish during the boiling? ____________
8. Set the dish on an asbestos pad to cool.
9. If there is any solid in the dish, carefully scrape it down into the liquid where it may be dissolved.
10. Pour the colored solution into a test tube. Compare it with the standard solution in the first test tube.
    (a) How do the colors of the two solutions compare? ____
    ____
    (b) Which solution seems to have more solute per milliliter of solution now? ____________________
    (c) Did reducing the amount of solvent dilute or concentrate the solution? ____________________

Some things to think about

1. How could you make a solution more concentrated without reducing the volume of the solution? __________
   ____
2. How would the density of a dilute solution compare with the density of a concentrated solution containing the same solute and solvent? ____________________
3. What is concentrated lemonade? ____________________
   How do you think it was made? ____________________
   What must you do before drinking the lemonade concentrate? ____________________


Does anybody know how to make fudge?
Is that part of the lab experiment?
I know why you asked that. Because fudge is a supersaturated sugar solution.
Super-saturated sugar solution!!? That's a mouth phull.

Saturation

In this experiment you will prepare unsaturated, saturated, and supersaturated solutions and find out how to identify them.

A SATURATED SOLUTION cannot hold any more solute at the existing temperature. The rate of dissolving equals the rate of crystallizing.

Experiment 6.4

Equipment and Supplies per Lab Station

beaker—250 ml; graduated cylinder—50 ml
(4) test tubes; balance; burner; forceps
test tube rack, test tube holder
30 g of sodium thiosulfate(hypo)
distilled water; crushed ice

What to do

1. Place 1 ml of water in a test tube. Add one crystal of hypo, which is sodium thiosulfate. Stopper and shake the test tube.
2. Find out whether the solution is unsaturated by trying to dissolve another crystal.

   Did the crystal dissolve? ______________________________

   Was the solution unsaturated? ______________________________
3. Pour 1 ml of water into a second test tube.
4. Add 1 g of hypo crystals. Stopper the test tube and shake it for about five minutes or until the crystals have dissolved.
   What did you observe besides the dissolving of the

   crystals? ______________________________

   How would you describe this change? ______________________________

5. Add another crystal to the solution made in step 4.

   Was the solution unsaturated or saturated? ________
6. Put 1 ml of water and 2 g of crystals in a third test tube. Stopper and shake the test tube for five minutes, unless all the crystals dissolve in less than five minutes.

   Is the solution formed saturated or unsaturated? ________

   How do you know this? ________
7. Add 1 ml of water and 25 g of hypo crystals to another test tube.

   Do you expect all of the crystals to dissolve? ________

   Why or why not? ________

   On the basis of observations made in step 4, how do you think you might be able to dissolve more crystals? ________

   ____
8. Heat the crystals and water gently until the mixture boils and a clear solution is formed. Your teacher will show you how to do this.
9. Cool the solution without shaking the test tube by holding the test tube under running water or by placing it in a beaker containing ice water.

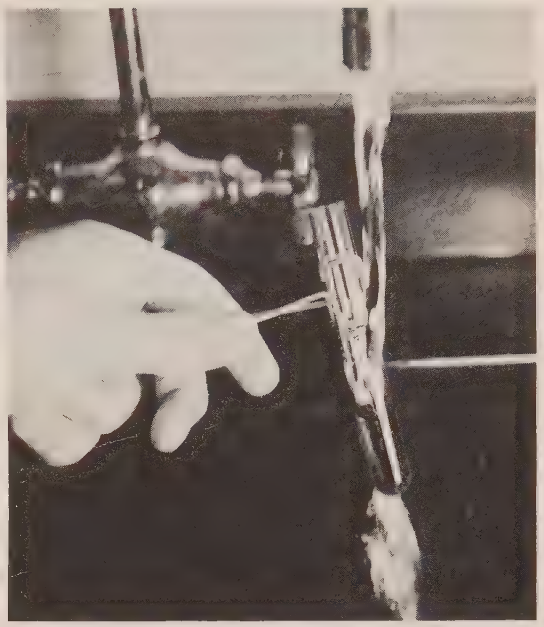
   Do crystals form during the cooling? ________

   What kind of a solution did you make? ________

   ____

   If the solution is unsaturated, what will happen when you add one more crystal? ________ If the solution is saturated? ________

   ____



10. Add the one more crystal to the solution from step 9 above. Write down all that you observe happening.

    (a) ________

    (b) ________

    Now, what kind of a solution do you think you have made? ________

    ____

Questions to answer

1. What do you think might happen if you stirred the supersaturated solution during the cooling process? Explain. ________

   ____
2. Which is more likely to happen as a hot saturated solution cools, crystallization or the formation of a supersaturated solution?

   ____
3. A crystal is dropped into a cold supersaturated solution. After the crystallization has stopped, what kind of a solution remains in the container? ________

   Explain. ________

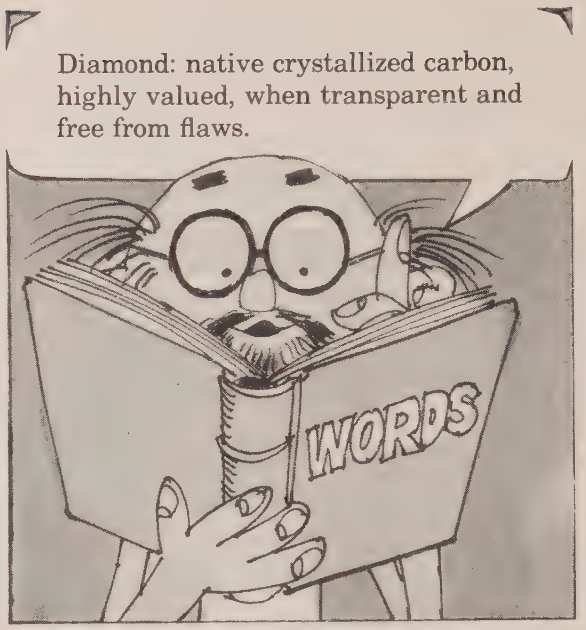
   ____

Extra for experts

When fudge is made, a supersaturated solution is formed before the candy is beaten. What would happen if some sugar crystals around the pan fell into the fudge as it was being beaten? ________

____

Diamonds are a girl's best friend.


Diamond: native crystallized carbon, highly valued, when transparent and free from flaws.
WORDS

Chemical Jewels

Crystals of chemicals, magnified so that even small ones can be seen, are very beautiful. Each chemical forms its own kind of crystal.

In this experiment you will make crystals and study their forms.

Experiment 6.4 – A Extension

Equipment and Supplies per Lab Station

beaker – 150 ml; graduated cylinder – 25 ml
glass stirring rod; watch glass; balance
magnifier; hand glass; microscope, if available
burner, ring stand with ring and asbestos wire gauze
waxed weighing papers
sample of the salt assigned; distilled water

What to do

First day

1. Obtain a sample of the chemical assigned to you. Your chemical is ______________________________
2. Find your chemical in the list below. Weigh on a waxed weighing paper the number of grams given for that chemical.

Sodium chloride	4 g
Nickel sulfate	34 g
Copper sulfate	20 g
Cobalt sulfate	8 g
Potassium chromate .	8 g
Cobalt chloride	19 g

3. Put the weighed chemical in a 150-ml beaker. Add 10 ml of distilled water. Stir the solution with a glass rod for a few seconds.
4. Warm the solution *slowly*, just to boiling. Stir all the while.
5. Turn out the burner. Leave the beaker on the ring stand until it is cool enough to handle.

6. Pour a few drops of the solution onto a watch glass. Cover the beaker with the watch glass. Transfer the beaker *carefully*, without shaking, to the shelf, or to a safe place as directed by your teacher. Let the covered beaker stand overnight or until the next laboratory period.
8. Small crystals may form on the watch glass almost right away. If so, scrape a *few* of these onto a glass plate. Examine them with a hand glass.

Second Day

1. Look at the crystals on the watch glass with a hand glass magnifier. Use a dark background if the crystals are white, or a white background if they are colored. Save the watch glass and crystals for comparison in step 3.
2. Decant the liquid from the beaker *without disturbing the solid.* How would you describe the solid? ________________

   ____
3. Shake a few crystals onto a filter paper. Transfer them, when they are dry, to a second watch glass. How do they compare with those on the first watch glass? ________________________
4. Examine the dry crystals from the beaker under the low power of the microscope, if possible. If not, use the hand magnifier. Record in your Data Chart:
   (a) The color of the crystals
   (b) A simple sketch to show the shape of the crystals
5. Examine the crystals in each different experiment in your class. Record in the Data Chart the color and a sketch of each different kind of crystal prepared.

DATA CHART

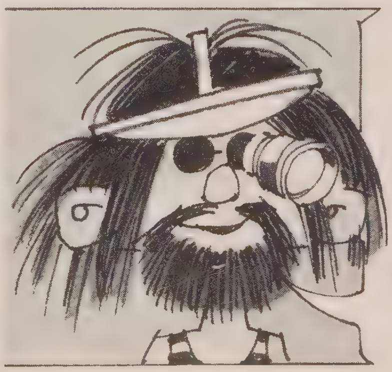
Chemical	Color	Sketch of Crystal
Sodium chloride		
Nickel sulfate		
Copper sulfate		
Cobalt sulfate		
Potassium chromate		
Cobalt chloride		

Questions to answer

1. How might a chemist make use of the fact that chemicals form crystals with characteristic forms? ________________

   ____
2. How do you think the beautiful jewels, such as amethyst, sapphire, and emerald may have been formed? ____________

   ____



Aww right, lady. Let's take a gander at that stone!

Vocabulary

DIRECTIONS: Write the letter "x" next to the term that best completes the statement or answers the question.

1. A true solution may be

(a) milky ____________ (c) colored ____________

(b) cloudy ____________ (d) opaque ____________

2. The part of the solution which does the dissolving is

(a) the solute ____________ (c) the precipitate ____________

(b) always water ____________ (d) the solvent ____________

3. The part of the solution which is dissolved is

(a) a vapor ____________ (c) the precipitate ____________

(b) the solute ____________ (d) the solvent ____________

4. A crystal of potassium nitrate is dissolved in a concentrated solution of potassium nitrate. The solution must have been

(a) unsaturated ____________ (c) saturated ____________

(b) supersaturated ____________ (d) dilute ____________

5. A solution that contains all the solute it can normally dissolve at a given temperature must be

(a) concentrated ____________ (c) supersaturated ____________

(b) saturated ____________ (d) unsaturated ____________

6. Which two liquids will dissolve in each other?

(a) ethanol and water ____________ (c) motor oil and rain water ____________

(b) salad oil and vinegar ____________ (d) hydrogen chloride and water ____________

7. Which procedure is most likely to increase the solubility of nearly all solids?

(a) stirring ____________ (c) heating the solution ____________

(b) pulverizing the solid ____________ (d) increasing the pressure ____________

8. When there is a great quantity of solid dissolved in a small amount of water, the solution is said to be

(a) saturated ____________ (c) dilute ____________

(b) supersaturated ____________ (d) concentrated ____________

9. A chemical of low solubility which drops out of solution is called

(a) a solute ____________ (c) a filtrate ____________

(b) a precipitate ____________ (d) a pulverized solid ____________

10. Two phase changes seen in distillation are

(a) evaporation and condensation ____________

(b) precipitation and decanting ____________

(c) filtration and deionization ____________

Laboratory Experiments in

ACTION CHEMISTRY

UNIT III

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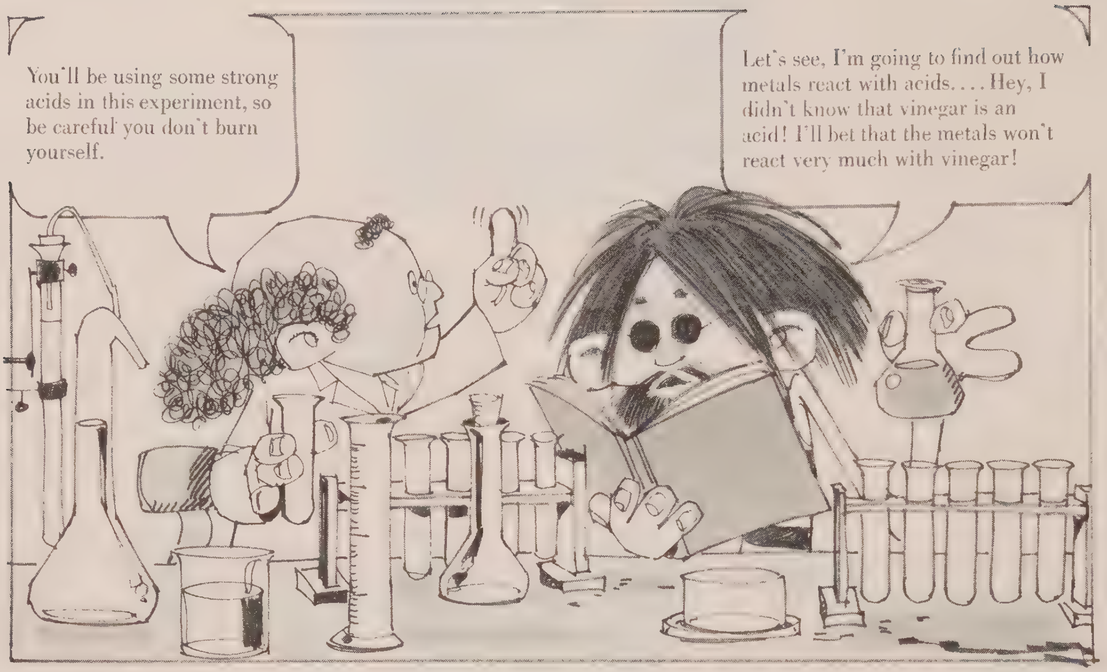
Chapter 7


You'll be using some strong acids in this experiment, so be careful you don't burn yourself.
Let's see, I'm going to find out how metals react with acids. . . . Hey, I didn't know that vinegar is an acid! I'll bet that the metals won't react very much with vinegar!

Metals and Acids

How does a metal react with acids? Does an acid react the same way with all metals? Look for the answers to these questions in this experiment.

Avoid accidents:

Put baking soda on any acid spilled on the desk before washing the acid up with wet sponges.

If you spill any acid on you, flush the area with lots of water. Tell the teacher if your skin appears to be burned or continues to itch.

Experiment 7.1

Equipment and Supplies per Lab Station

beaker-250 ml; forceps; (4) test tubes; test tube rack

If hot water is not available from the tap, the following are also needed: asbestos wire gauze; burner; ring stand with ring; aluminum foil strip; baking soda (for spilled acid) copper; iron; (4) magnesium strips about 1-inch long

steel wool; zinc;

Solutions of: acetic acid; hydrochloric acid; sulfuric acid; vinegar

If Extra for Experts is done, the following are also needed: matches; wooden splints

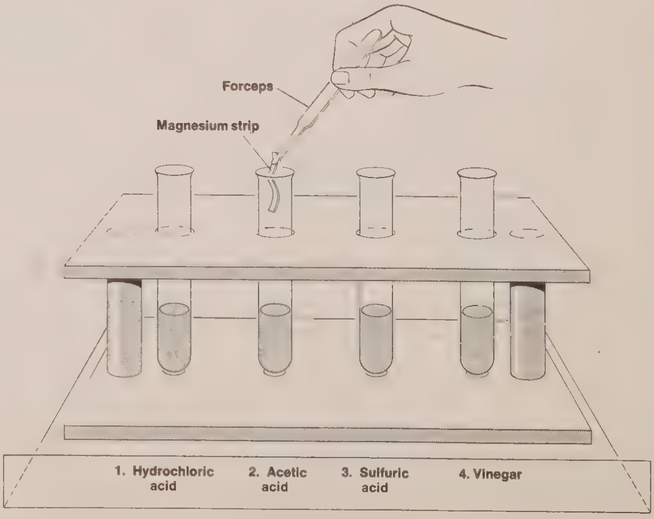


Fig. 28.

What to do

PART I

1. Obtain 8 to 10 ml of dilute solutions of hydrochloric, acetic, and sulfuric acids, and of vinegar in four separate labeled test tubes. See Fig. 28.
2. If the magnesium is dull, clean 4 one-inch pieces of magnesium by rubbing them with steel wool.
3. Add a piece of magnesium to each test tube. Record what happens in Data Chart 1.
4. Place any test tube in which a reaction seems slow in starting into a beaker of hot water until the reaction begins.

DATA CHART I

Test Tube Number	Solution	Formula	Reaction with magnesium	Test for gas (optional)
1	Hydrochloric acid	HCl		
2	Acetic acid	$HC_2H_3O_2$		
3	Sulfuric acid	H_2SO_4		
4	Vinegar			

PART II

1. Pour 5 ml of hydrochloric acid into each of four test tubes.
2. Obtain small pieces of zinc, copper, aluminum, and iron.
3. Add a different metal to the acid in each test tube. Write your observations in Data Chart 2.

DATA CHART 2

Metal	Symbol	Reaction with hydrochloric acid	Test for gas (optional)
Zinc	Zn		
Copper	Cu		
Aluminum	Al		
Iron	Fe		

Questions to answer

1. What answers have you found to the questions asked at the beginning of this experiment?

____

2. How can you account for the way vinegar behaves? _______________

____

3. What chemical element is found in the formula of all of the three laboratory acids? _______________

Extra for experts

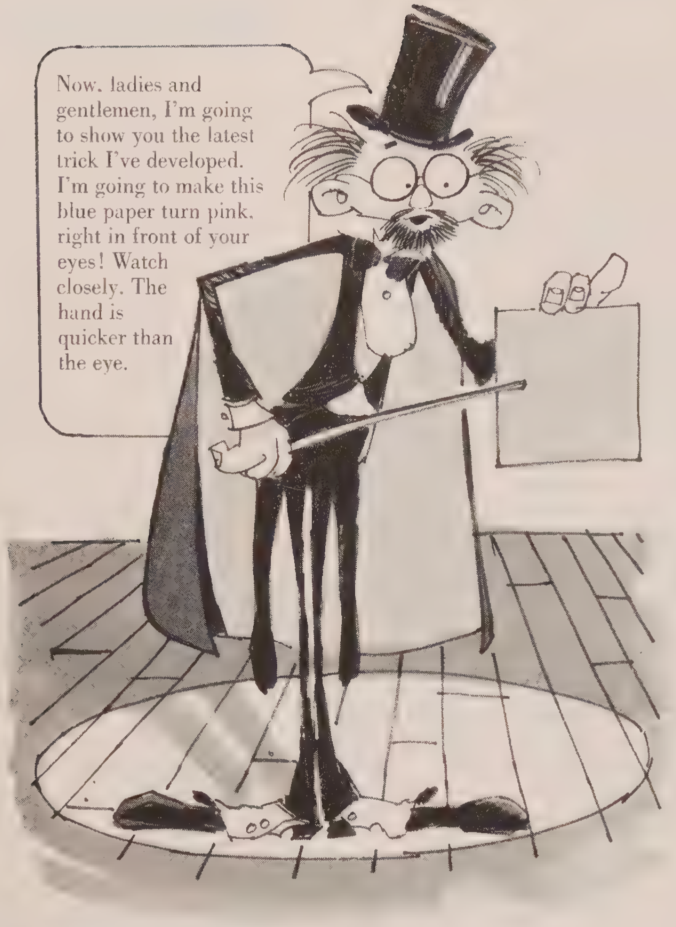
1. Look at How to Use a Splint to Identify Some Gases on page *3* / 18. If your teacher approves, test any gas given off in any of the reactions in Parts I and II.

   (a) What test did you use? _______________

   ____

   (b) What gas did you find present? _______________

   ____


Now, ladies and gentlemen, I'm going to show you the latest trick I've developed. I'm going to make this blue paper turn pink, right in front of your eyes! Watch closely. The hand is quicker than the eye.

Acids Turn the Color of Indicators

A compound which reacts with acids to show a definite color change is an acid INDICATOR. A chemist finds indicators very useful. In a matter of seconds he can find out whether a liquid behaves as an acid. Which indicators can you use to identify acid solutions?

Experiment 7.2

Equipment and Supplies per Lab Station
glass plate; stirring rod; (3) test tubes; test tube rack
white paper; litmus paper, red and blue
Solutions of: acetic acid; bromphenyl blue indicator
hydrochloric acid; methyl orange indicator; sulfuric acid;
unknowns

What to do

PART I

1. Lay a clean dry watch glass or glass plate on a piece of white paper.
2. Place a strip of blue litmus paper and red litmus paper on the plate.
3. Pour 5 ml of hydrochloric acid into each of three test tubes.
4. Dip a clean, dry, stirring rod into the acid in the first test tube. Then touch the stirring rod to the litmus paper. Write your observations in Data Chart 1.
5. Rinse and dry the stirring rod.
6. Repeat steps 4 and 5 using the red litmus paper.

7. Add two drops of bromphenyl blue solution to the acid in the second test tube. Record your observations in Data Chart 1.
8. Add two drops of methyl orange to the acid in the third test tube. Record your observations in Data Chart 1.
9. Repeat steps 1–8 using sulfuric acid in place of hydrochloric acid.
10. Repeat steps 1–8 using acetic acid instead of hydrochloric acid.

DATA CHART 1

INDICATOR	Indicator Color Change in Unknown		
	Hydrochloric Acid	Sulfuric Acid	Acetic Acid
Red Litmus			
Blue Litmus			
Bromphenyl Blue			
Methyl Orange			

PART II

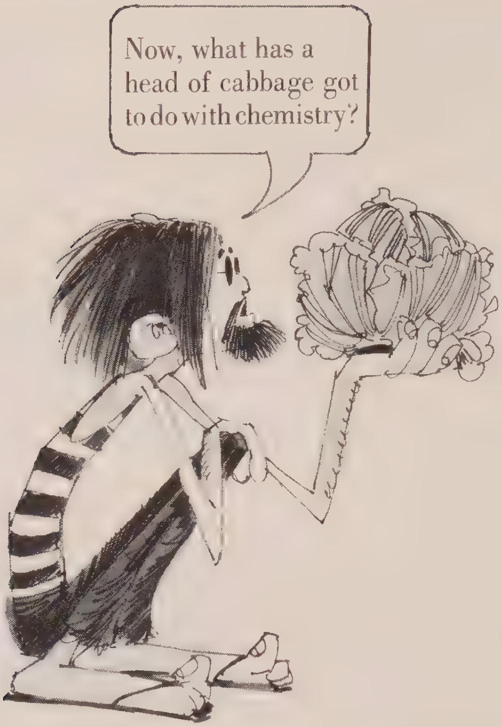
1. Obtain a set of unknown solutions. Use the indicator tests to determine which solutions are acids. Record your observations in Data Chart 2.

DATA CHART 2

INDICATOR	Indicator Color Change in Unknown			
	1	2	3	4
Red Litmus				
Blue Litmus				
Bromphenyl Blue				
Methyl Orange				

2. Study the Data Chart 2. Which unknown solutions were acids? ______________________
3. Check with your teacher to see if you were right. How many did you get right? __________

   ____
4. Save the unknown solutions to use in Experiment 7.3.


Now, what has a head of cabbage got to do with chemistry?

Making an Indicator from Cabbage

Some indicators may be prepared from plants. One indicator is extracted from the leaves of red cabbage.

In this experiment you are going to prepare a cabbage indicator and use it to test acidity. You may even be able to identify the indicator by another name!

Experiment 7.2-A Extension

Equipment and Supplies per Lab Station

(2) beakers-250 ml; (3) test tubes; asbestos wire gauze burner; ring stand with ring; test tube rack
(3) stoppers to fit test tubes; red cabbage (purple); water, distilled or deionized
Solutions of: acetic acid; hydrochloric acid; sulfuric acid

What to do

1. Tear a leaf of red (purple) cabbage into small pieces. Pack the pieces tightly against the bottom of a 250-ml beaker.
2. Cover the pieces of leaf with distilled water. Use as little water as possible.
3. Boil the cabbage until the liquid is a deep purple color.
4. Cool the contents of the beaker, then decant the colored liquid into another beaker.
5. Pour 3 or 4 ml of hydrochloric acid into a test tube. Add enough cabbage indicator to form a deep color.

6. Stopper the test tube and shake it to mix the two liquids. Record any observations in the Data Chart.
7. Repeat steps 5 and 6 using sulfuric acid. Repeat steps 5 and 6 using acetic acid. In both cases record your observations in the Data Chart.

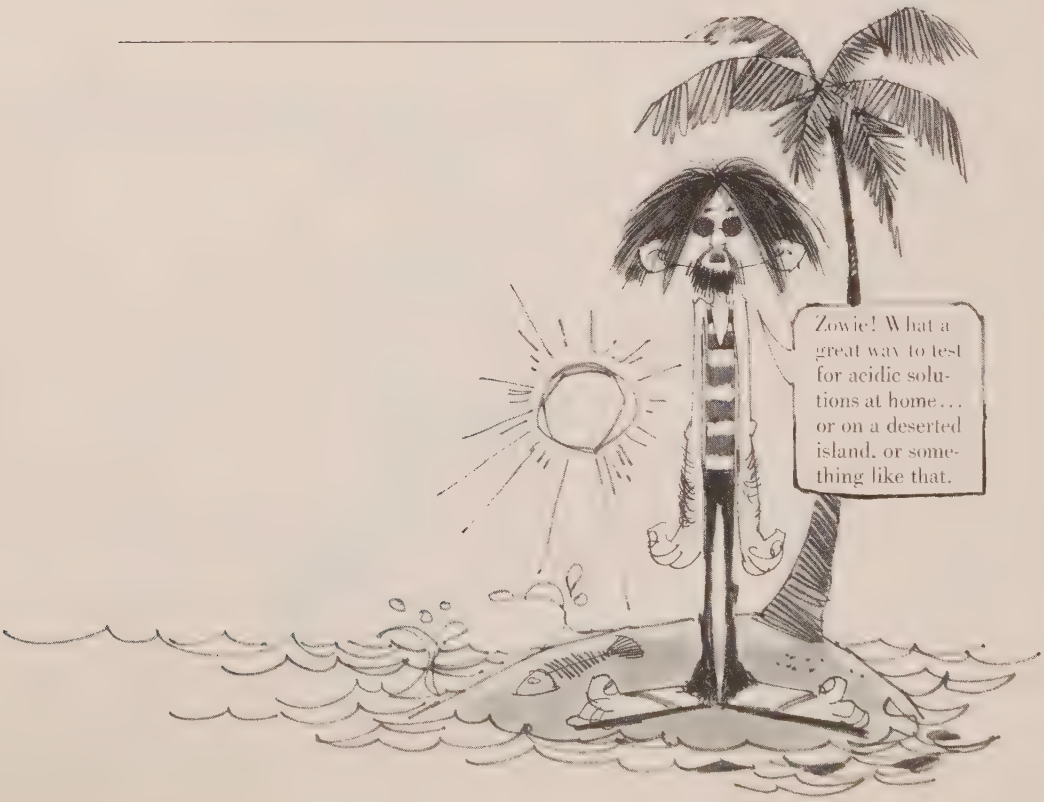
DATA CHART

Color Cabbage Indicator	Acid Used	Color Cabbage Indicator In Acid
	hydrochloric acid	
	sulfuric acid	
	acetic acid	

What color is cabbage indicator in the presence of acids? ______________________________

Questions to answer

1. What do you think is the other name for the cabbage indicator? ____________ Explain.
____
____
2. Why does purple cabbage turn red when vinegar is added to it during the cooking process? ______________________________
____
____

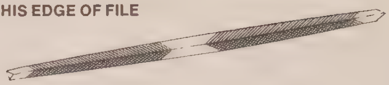

Zowie! What a great way to test for acidic solutions at home... or on a deserted island, or something like that.

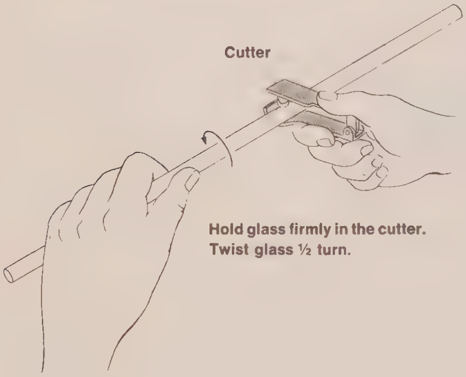
HOW TO

make a dropper

1. Make a deep scratch across the glass tubing.
   Use a file or a glass tubing cutter.

USE THIS EDGE OF FILE

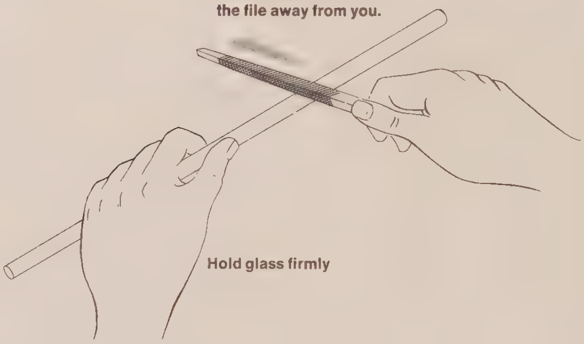




Cutter

Hold glass firmly in the cutter.
Twist glass ½ turn.

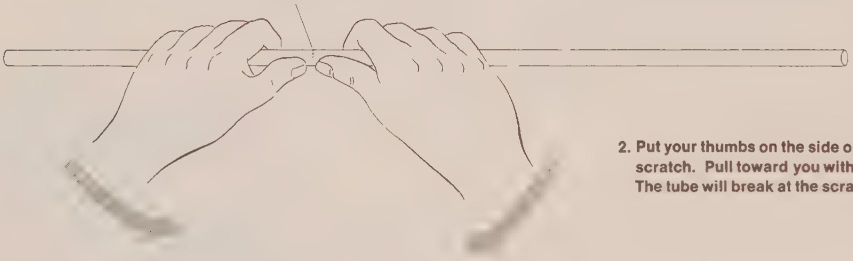
OR



Push down and move
the file away from you.

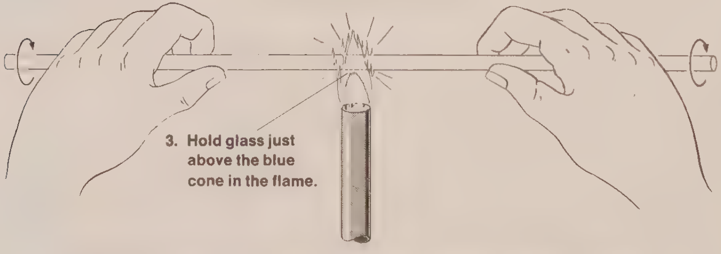
Hold glass firmly

SCRATCH SHOULD BE
AWAY FROM YOU.



2. Put your thumbs on the side opposite the
   scratch. Pull toward you with BOTH hands.
   The tube will break at the scratch mark.

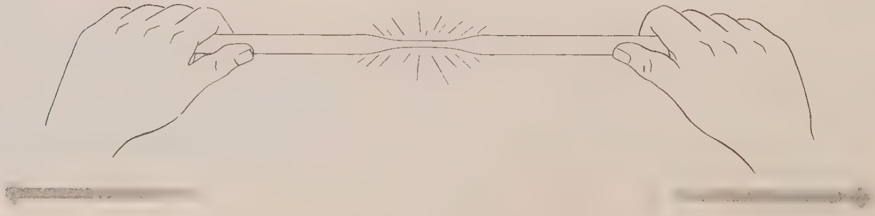
HEAT IN SAME PLACE.



Keep rotating the tube
until the glass gets soft
and tends to bend.

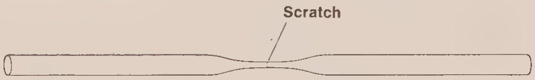
3. Hold glass just
   above the blue
   cone in the flame.

4. Remove the glass from the flame. PULL out to "stretch" the glass a little bit.

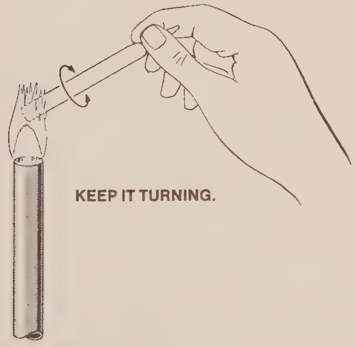


YEEEEOOOWW!
Hot glass doesn't look hot, but it'll give you a good burn. Hey Moony, make that noise again!

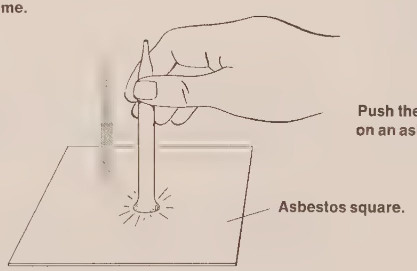
5. Carefully make a scratch in the narrow part of the tubing. Break the tube as you did in step 2.



6. Heat the larger end in the flame.



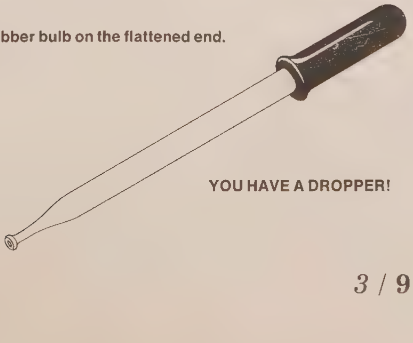
7. When the flame turns yellow and the glass begins to get soft, remove it from the flame.



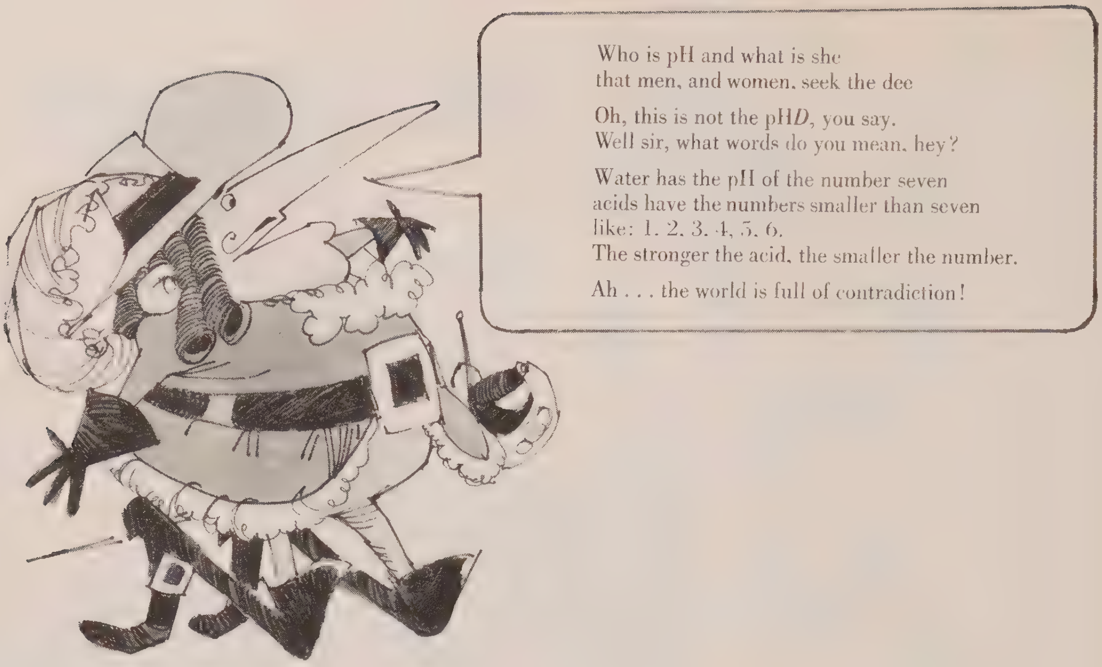
Push the hot end down slightly on an asbestos square to flatten it.

8. When the flattened end has cooled repeat step 6 for the narrow end. A little heat will do.

9. Put the rubber bulb on the flattened end.



YOU HAVE A DROPPER!


Who is pH and what is she
that men, and women, seek the dee
Oh, this is not the pH*D*, you say.
Well sir, what words do you mean, hey?
Water has the pH of the number seven
acids have the numbers smaller than seven
like: 1, 2, 3, 4, 5, 6.
The stronger the acid, the smaller the number.
Ah . . . the world is full of contradiction!

How Strong Is It?

Chemists have invented a system of measuring how acid a substance is. The system is known as pH or hydrogen potential. pH is based upon water which is not acid. The pH of water is said to be 7. A substance which is more acid than water has a pH of less than 7. Acids have a pH range from 0 to just below 7.

In this experiment we will use the label *M* to indicate concentration. We will learn more about this symbol in Unit IV.

How does the concentration of an acid affect the pH?

Experiment 7.3

Equipment and Supplies per Lab Station

dropper with long stem*; (2) flasks—150 ml or 250 ml
(2) graduated cylinders—10 ml; (3) test tubes
watch glass or a glass plate; (2) stoppers to fit flasks
test tube rack; labels or a marking pen; sheet of white paper
universal indicator paper; water, distilled or deionized
Solutions of: hydrochloric acid; methyl indicator
*If you make the dropper, the following are also needed:
asbestos pad; burner; a rubber bulb
glass tubing cutter or file; glass tubing about 12 in long

What to do

DO NOT DISCARD ANY ACID UNTIL INSTRUCTED TO DO SO.

1. Pour 99 ml of distilled or deionized water into a 250-ml flask.
2. In a 10-ml graduated cylinder measure 10 ml of 0.1 *M* hydrochloric acid.

3. Use your dropper to transfer 1 ml of the acid a drop at a time to the flask with the water in it.
4. Put a stopper in the flask. Gently shake the flask to mix the two liquids. ALWAYS ADD ACID TO WATER! NEVER ADD WATER TO ACID.
5. Label the flask 0.001*M* acid.
6. Pour two or three ml of the 0.001*M* acid you just made into a test tube. Rinse the dropper several times with this acid.
7. Repeat steps 1–5 with the 0.001*M* acid you made.
8. Label the new solution 0.00001*M* acid.
9. Arrange the three samples of acid in the order of most concentrated to the least concentrated.

   Which is the most concentrated? ______________

   Explain. ______________________________

   ____
10. Place 3 strips of universal indicator paper on a glass plate or watch glass. Place the glass on a sheet of white paper.
11. Using a glass stirring rod transfer a drop of the 0.1*M* acid to one end of one of the indicator strips. Label the strip with the concentration of the acid.
12. You will find a color key on the universal indicator's container. The colors shown are the "pH" with a number. Compare the color of the strip with the key. Write the corresponding pH number in the Data Chart.
13. Repeat steps 10–12 using each of the other two acid concentrations. Record all your observations in the Data Chart.
14. Pour 5 ml of each concentration of the acid into separate test tubes. Label each tube with the concentration of the acid in it. Add 2 drops of methyl orange indicator to each sample. Include your observations in the Data Chart.

DATA CHART

Concentration of the acid	Universal Indicator		Color with methyl orange
	color	pH	
0.1*M*			
0.001*M*			
0.00001*M*			

15. Discard the acid samples.

Questions to answer

1. How do the color changes with the methyl orange compare with those with the universal indicator? ______________________________
2. Would you use methyl orange to find how strong an acid is? Explain. ______________

   ____

   ____

   ____

   ____

Extra for experts

Test the unknown solutions from Experiment 7.2.

What pH do you find for each? ______________

____

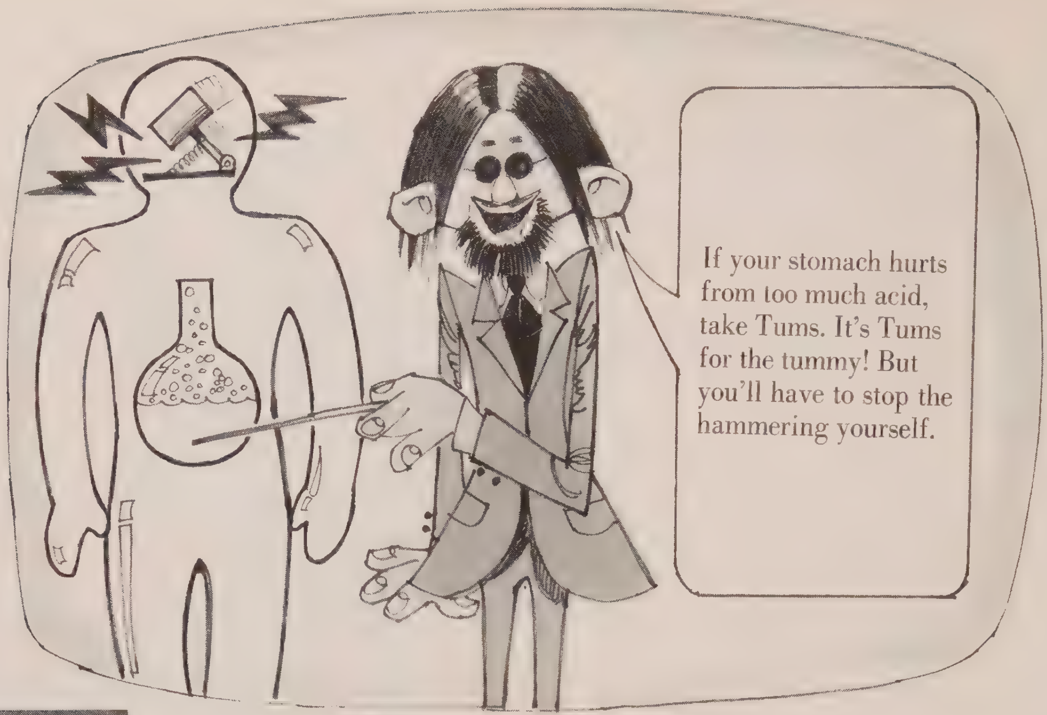
____

____

Check with your teacher to find out if you are right. How many did you get right? ______________

____

____


If your stomach hurts from too much acid, take Tums. It's Tums for the tummy! But you'll have to stop the hammering yourself.

Experiment 7.4

THE BUBBLERS

On the supply table there are a number of powders and tablets which you have seen advertized on TV, found in the supermarket, or used in your home. How are these substances alike? Just what are the bubblers?

Equipment and Supplies per Lab Station
dropper; gas bottle or flask – 250 ml; glass plate
graduated cylinder; spoon; (6) test tubes
test tube rack; wooden splints
baking soda (sodium hydrogen carbonate); citric acid crystals
cream of tartar (potassium hydrogen tartrate)
calcium hydrogen phosphate; litmus, both red and blue
substances to be tested – commercial items; tartaric acid

PART I

What to do

1. Put ½ tablet or 1 teaspoonful of one of the substances in a gas bottle or Erlenmeyer flask.
2. Add 50 ml of water to the substance.
3. Carefully insert a burning splint part way into the bottle. Record all your observations in Data Chart 1. See Fig. 29, and page *3* / 18.
4. Repeat steps 1–3 using a different substance each time.

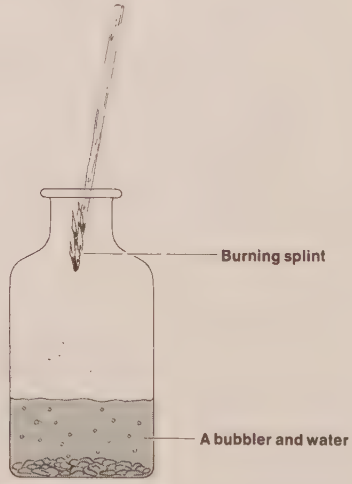


Fig. 29.

DATA CHART 1

Substance tested	What happens when	
	water is added	burning splint is used

PART II

What to do

1. Put a small amount of each chemical in a test tube or on a glass plate. Label each. See Fig. 30.
2. Add 5 drops of water to each substance. Test each solution with red and with blue litmus. Record all observations in Data Chart 2.
3. Mix ½ spoonful of one powder with ½ spoonful of another substance. Put the mixture in a test tube.
4. Add a few drops of water to the mixture. Test any gas with a burning splint. Write all observations in Data Chart 3.
5. Repeat steps 3 and 4 using different combinations of the substances. Write all your data in the chart.

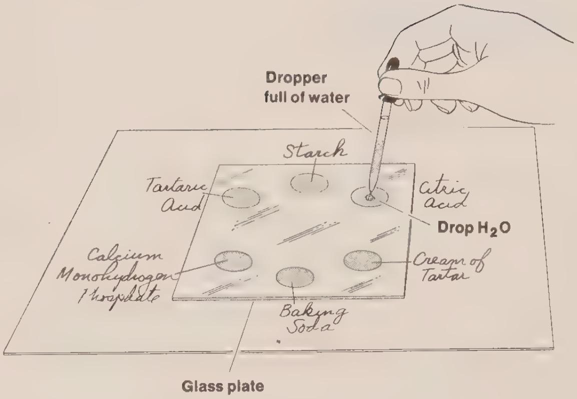


Fig. 30.

DATA CHART 2

Number	Name of substance	Effect on litmus		A bubbler?
		red	blue	
1.				
2.				
3.				
4.				
5.				

DATA CHART 3

Mixture		Bubbles with water?	Burning splint test
Substance 1	Substance 2		

Questions to answer

1. What product is formed by the bubblers? (Burning splint test can identify the gas.)
2. What two kinds of substances seem to make up a bubbler combination?
3. What substances do you think were present in the materials used in PART 1?

Check your answers with the contents listed on the labels.

Exercise

Using Acids in Word Equations

When acids react with other substances, they form a salt and at least one other substance. For example:

(a) a salt + an acid → a new salt + a new acid
(b) a carbonate + an acid → a salt + water + carbon dioxide
(c) a metal + an acid → a salt + hydrogen

The reference table, ACIDS AND THE COMPOUNDS THEY FORM (found in the APPENDIX), helps you learn the names of the salts that acids make. The first part of any salt's name comes from the other reacting substance. The last part of the salt's name comes from the acid used.

$$\text{zinc} + \text{hydrochloric acid} \rightarrow \overbrace{\underbrace{\text{zinc chloride} + \text{hydrogen}}_{\text{last part from acid}}}^{\text{first part from other reactant}}$$

DIRECTIONS: Complete the word equation:

1. magnesium + hydrochloric acid → ____________________ + ____________________

2. zinc + sulfuric acid → ____________________ + ____________________

3. aluminum + acetic acid → ____________________ + ____________________

4. barium + sulfuric acid → ____________________ + ____________________

5. hydrochloric acid + aluminum → ____________________ + ____________________

6. hydrochloric acid + iron → ____________________ + ____________________

7. magnesium + phosphoric acid → ____________________ + ____________________

8. aluminum + sulfuric acid → ____________________ + ____________________

8. zinc + acetic acid → ____________________ + ____________________

9. sodium carbonate + hydrochloric acid → ______________ + ______________ + ______________

10. calcium carbonate + hydrochloric acid → ______________ + ______________ + ______________

11. magnesium carbonate + sulfuric acid → ________ + ________ + ________

12. aluminum carbonate + nitric acid → ________ + ________ + ________

13. zinc carbonate + acetic acid → ________ + ________ + ________

14. sodium carbonate + nitric acid → ________ + ________ + ________

15. calcium carbonate + nitric acid → ________ + ________ + ________

EQUATION REVIEW

Complete the following word equations:

16. copper + oxygen → ________ + ________

17. zinc + sulfur → ________ + ________

18. aluminum + chlorine → ________ + ________

19. hydrogen + oxygen → ________ + ________

20. sodium chloride + silver nitrate → ________ + ________

21. aluminum sulfate + barium chloride → ________ + ________

22. magnesium nitrate + sodium carbonate → ________ + ________

23. sodium chloride + lead nitrate → ________ + ________

24. hydrochloric acid + potassium hydroxide → ________ + ________

25. potassium carbonate + sulfuric acid → ________ + ________ + ________

Chapter 8

Let's see what's in the first base.
I'll set up the second base of the experiment.
I didn't know we were going to play baseball in chem lab?

A Detective Assignment

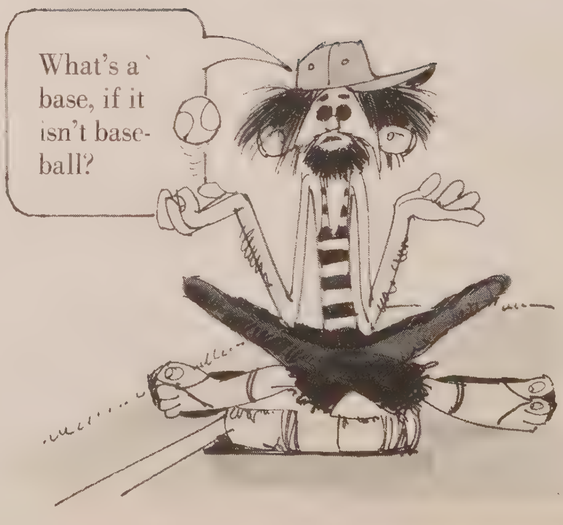
Household ammonia is used for tough cleaning jobs around the house. Does this cleaner belong to the "gang" known as bases?

Investigate some properties of bases and see whether household ammonia behaves like bases do.

Experiment 8.1

Equipment and Supplies per Lab Station

glass plate or watch glass; stirring rod; (4) test tubes
test tube rack; white paper; litmus paper – red and blue
universal indicator paper;
Solutions of: ammonia water (household ammonia); methyl orange indicator
phenolphthalein indicator; potassium hydroxide or barium hydroxide; sodium hydroxide


What's a base, if it isn't baseball?

What to do

KEEP THE CONTAINERS CLOSED WHEN NOT IN USE.

1. Pour about 5 ml of dilute sodium hydroxide solution into a test tube.
2. Place a clean glass plate or a watch glass on a sheet of white paper.
3. Lay a strip of each of the following on the glass: red litmus, blue litmus, and universal indicator.
4. With a clean stirring rod, transfer a drop of the sodium hydroxide onto each indicator. Be sure to wash and dry the stirring rod after touching an indicator paper. Record your observations in the Data Chart.

5. Compare the color of the universal indicator with the pH-color chart. Write down the pH of the base.
6. Pour half of the remaining sodium hydroxide into another test tube.
7. To one sample of the base add 2 drops of phenolphthalein. Record your observations in the Data Chart.
8. Add 2 drops of methyl orange into the other sample of the base. Record your observations in the Data Chart.
9. Cautiously note the odor of the base. See Fig. 31.

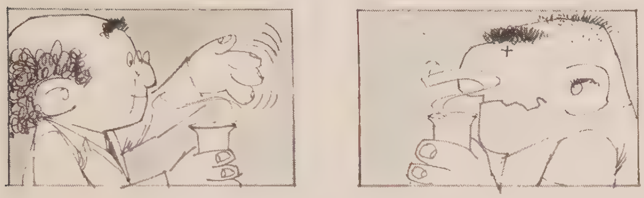


Fig. 31. Do it this way, not this way.

10. Place 1 drop of the sodium hydroxide in a test tube. Add 9 drops of water. Mix well.
11. With a stirring rod transfer a drop of the diluted base to your thumb. Rub your thumb and a finger together. What do you feel?

    ______________ Wash the base off your skin. (ALWAYS WASH OFF ANY BASE WHICH YOU GET ON YOUR SKIN FOR BASES CAN CAUSE BURNS.)
12. Repeat steps 1–11 using any one of the other hydroxide compounds on the supply table. Record your observations in the Data Chart.
13. Repeat steps 1–11 using the household ammonia solution. Record your observations in the Data Chart.
14. Place a drop of sodium hydroxide on a watch glass. On another watch glass put a drop of household ammonia. Set the glasses under a hood until the next class period. Then inspect them. Is there any residue left on the watch glasses? ______________________________

    ____

    ____

DATA CHART

Indicator	Sodium Hydroxide	Another Hydroxide (name:)	Household Ammonia
Red Litmus			
Blue Litmus			
Universal Indicator Color			
pH from Color Chart			
Phenolphthalein			
Methyl Orange			
Feel of Solution			
Odor			

Questions to answer

1. What properties do bases have?

   (a) ______________ (c) ______________

   (b) ______________ (d) ______________
2. Would you classify household ammonia as a base?______________ Explain______________

   ____
3. How can you tell from its name whether a substance is base? ______________________________

   ____
4. Refer to your Data Chart for Experiment 7.2. How do bases change the colors of indicators differently from acids? ______________________________

   ____

Extra for experts

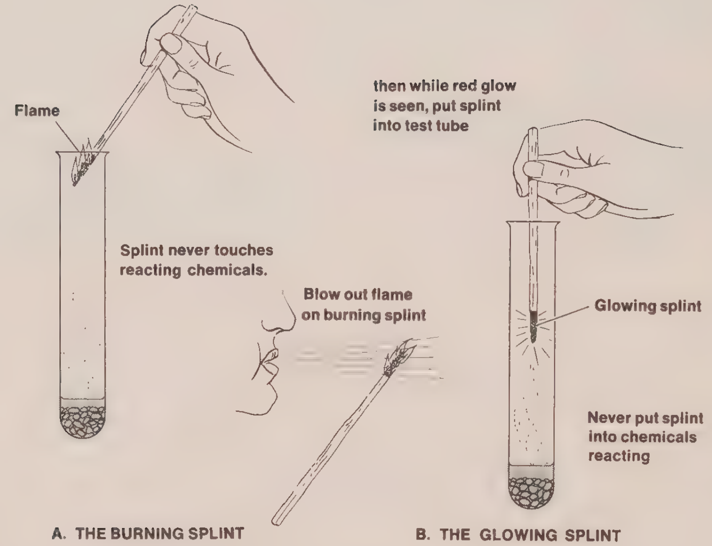
Why is household ammonia used instead of sodium hydroxide in cleaning windows?

____

____

HOW TO

use a splint to identify some gases


Flame
Splint never touches reacting chemicals.
Blow out flame on burning splint
then while red glow is seen, put splint into test tube
Glowing splint
Never put splint into chemicals reacting

A. THE BURNING SPLINT

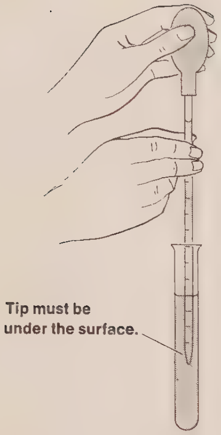
B. THE GLOWING SPLINT

HOW TO USE A SPLINT TO IDENTIFY SOME GASES

WHAT TO USE	IF THE GAS IS	OBSERVE
A burning splint	carbon dioxide	Flame goes out.
A burning splint	hydrogen	Bark-like noise is heard; flame may flash back into container.
A glowing splint	oxygen	Splint bursts into flame.

HOW TO

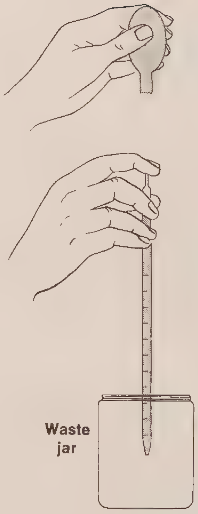
use a Pipet


Tip must be under the surface.

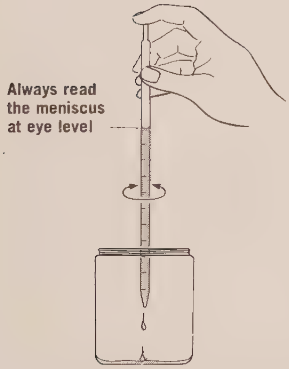
A pipet has been made to measure volumes of liquid more accurately than a graduated cylinder.

Step 1: Place a bulb on the top of a clean pipet. (Your teacher will show you how to clean a pipet). Squeeze the bulb, then put the tip of the pipet into the liquid to be measured. Release the bulb gradually until the liquid rises above the zero mark. *AVOID GETTING THE LIQUID INTO THE BULB!*

Don't put pipet into reagent bottle. Transfer some liquid to a test tube.

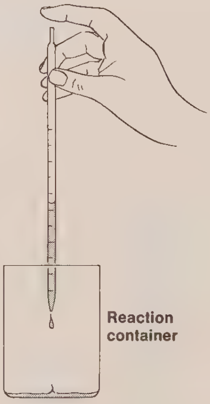


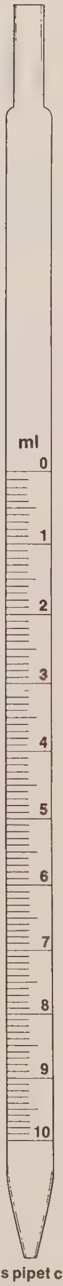
Step 2: Hold the bulb and pipet over a waste container. As you remove the bulb, quickly slide your finger over the top of the pipet, and press down. Pressing down on the top of the pipet keeps the liquid from dripping out of the pipet.


Always read the meniscus at eye level

Step 3: Roll the pipet back and forth between your fingers until the meniscus is level with the zero mark. Hold your finger on the top to regulate the flow of liquid.

Step 4: Hold the pipet over the reaction container. Lift your finger, and let the liquid flow out until the meniscus is level with the mark for the volume you wish to deliver. Then quickly press your finger down on the pipet's top. Discard the liquid remaining into the waste jar.





This pipet can deliver from 0.1 ml to 10.0 ml.

My love for you is stronger than dirt!

What Is Its Strength?

Chemists use indicators to test the strength of solutions of bases just as they do for acid solutions. The colors may be different, however. Try it yourself!

Experiment 8.1-A Extension

Equipment and Supplies per Lab Station

graduated cylinder—10 or 25 ml; pipet—0-10.0 ml range
stirring rod; (4) test tubes
bulb for pipet; stoppers for test tubes; test tube rack
labels or a marking pen; water, distilled or deionized
universal indicator test paper with color pH scale on container
Solutions of: indigo carmine indicator; sodium hydroxide

What to do

1. Use a pipet to put 10 ml of 0.1*M* sodium hydroxide in a test tube. Label the test tube 0.1*M* sodium hydroxide.
2. Place 9 ml of distilled water in each of three other test tubes. Use a clean pipet to measure the water.
3. With a clean pipet, transfer 1 ml of 0.1*M* sodium hydroxide solution to one of the test tubes. See Fig. 32.
4. Put a stopper in a test tube, and gently shake the tube to mix the liquids.
5. Label the solution you just made 0.01*M* sodium hydroxide. It is 0.1 as strong as the base you used.
6. Repeat steps 3–5, but use the 0.01*M* sodium hydroxide solution you just made. Label the new solution 0.001*M* sodium hydroxide.

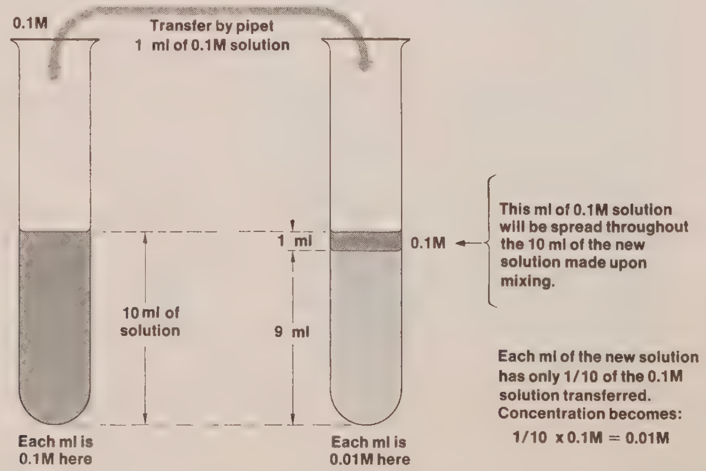

0.1M
Transfer by pipet 1 ml of 0.1M solution
1 ml
0.1M
This ml of 0.1M solution will be spread throughout the 10 ml of the new solution made upon mixing.
10 ml of solution
9 ml
Each ml of the new solution has only 1/10 of the 0.1M solution transferred. Concentration becomes:
1/10 x 0.1M = 0.01M
Each ml is 0.1M here
Each ml is 0.01M here

Fig. 32.

7. Repeat steps 3–5 using the 0.001*M* solution. Label the new solution you make 0.0001*M* sodium hydroxide.
8. Use a clean stirring rod to transfer a drop of 0.1*M* solution to a strip of universal test paper. Immediately compare the color with the color chart on the vial. Write all your observations and the pH reading in the Data Chart.
9. Rinse and dry the stirring rod. Repeat step 8 for each of the other solutions. Record all data in the chart.
10. To each of the tubes of sodium hydroxide add two drops of indigo carmine indicator. Stopper each test tube, and gently shake it to mix the liquids. Record any observations in the Data Chart.

DATA CHART

Concentration of the base	Universal Indicator		Color with indigo carmine
	color	pH	
0.1*M*			
0.01*M*			
0.001*M*			
0.0001*M*			

Questions to answer

1. Would a solution having a pH of 12 or 13 be a strong or weak acid or base? ______________

   Explain. ______________________________

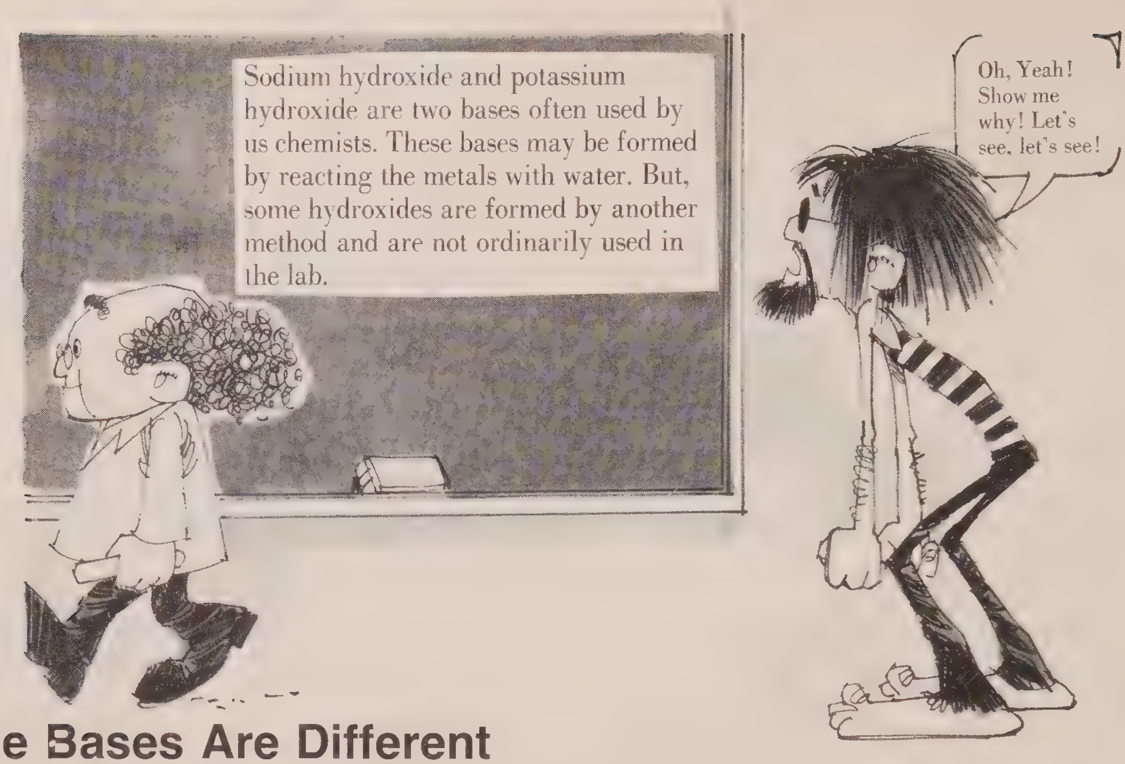
   ____

2. Can indigo carmine indicator be used to determine pH? ______________________

   Why? ______________________________

Extra for experts

If your teacher approves, find the concentration of the sodium hydroxide solution used in Experiment 8.1 by comparing its pH with the dilutions in the Data Chart.


Sodium hydroxide and potassium hydroxide are two bases often used by us chemists. These bases may be formed by reacting the metals with water. But, some hydroxides are formed by another method and are not ordinarily used in the lab.
Oh, Yeah! Show me why! Let's see, let's see!

Some Bases Are Different

Sodium hydroxide and potassium hydroxide are two bases often used by chemists. These bases may be formed by reacting the metals with water.

Since sodium and potassium react with air, you will find them always stored in kerosene.

Experiment 8.2

Equipment and Supplies per Lab Station

(2) beakers—250 ml; dropper; funnel
graduated cylinder—100 ml; stirring rod; (8) test tubes
forceps; (4) stoppers to fit test tubes; test tube rack
filter paper; litmus paper, red; potassium metal
potassium hydroxide pellets; phenolphthalein; sodium metal,
sodium hydroxide pellets;
water, distilled or deionized
Solutions of: cobalt (II) nitrate; copper (II) nitrate;
iron (III) nitrate; zinc nitrate

What to do:

PART I TEACHER DEMONSTRATION

1. Fill a 250-ml beaker about $^{2}/_{3}$ full of distilled water. Add two drops of colorless phenolphthalein indicator. What color is the indicator in the water? ______________
2. Cautiously drop in a piece of sodium metal about the size of the head of a large pin. Use forceps. Stand back until the reaction has been completed.

   Describe what happens. ______________

   ____

   What evidence is there that sodium hydroxide has been formed by the reaction? ______________
3. Inspect a bottle of sodium hydroxide solid. Read the precautions to be followed on label.

   Describe its appearance. ______________
4. Using your forceps, put one pellet of sodium hydroxide into a beaker containing 100 ml of water. Stir. What do you observe? ______________
5. Add two drops of colorless phenolphthalein indicator to the liquid in the beaker. What does the indicator's color show? ______________
6. If your teacher approves, repeat steps 1–6 substituting potassium and potassium hydroxide for the sodium and sodium hydroxide.

In Part II some bases are made by another method. Why can't sodium and potassium hydroxides be made by this other method?

PART II

1. Transfer 5 ml of copper (II) nitrate solution into a test tube.
2. Add 5 ml of sodium hydroxide solution. Copper (II) hydroxide is formed. Write your observations in the Data Chart.
3. Filter the contents of the test tube. Discard the filtrate.
4. Transfer the funnel so the stem is in a dry test tube. Poke a hole in the tip of the paper in the funnel. Wash the material off the filter paper with 5 ml of distilled water. See Fig. 33.
5. Put a stopper in the test tube. Shake it to mix the materials.
6. Use your forceps to hold a piece of red litmus. Dip the litmus into the liquid in the test tube. Add your observations to the Data Chart.
7. Add 5 ml of sodium hydroxide drop by drop to the contents of the test tube. Record your observations in the Data Chart.
8. Repeat steps 1–7 with each of the following in place of the copper (II) nitrate: iron (III) nitrate, cobalt (II) nitrate, and zinc nitrate. Complete the Data Chart with your observations. In each case a hydroxide compound is made.

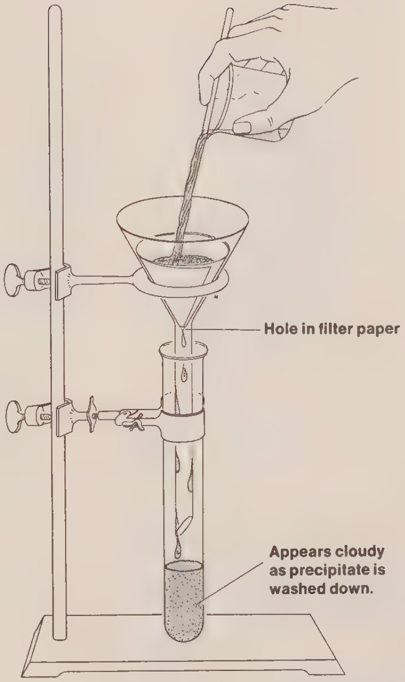


Fig. 33.

DATA CHART

Hydroxide formed	Color	Soluble	Litmus test	Effect of extra NaOH
copper(II) hydroxide				
iron(III) hydroxide				
cobalt(II) hydroxide				
zinc hydroxide				
sodium hydroxide				xxxxxxxxxxx
potassium hydroxide				xxxxxxx

Questions to answer

1. What color do bases turn litmus? ______________
   phenolphtalein? ______________________________
2. Are most hydroxides soluble or insoluble? (The SOLUBILITY REFERENCE TABLE in the APPENDIX of Unit II lists the solubility of many bases.) ______________________________

Extra for experts

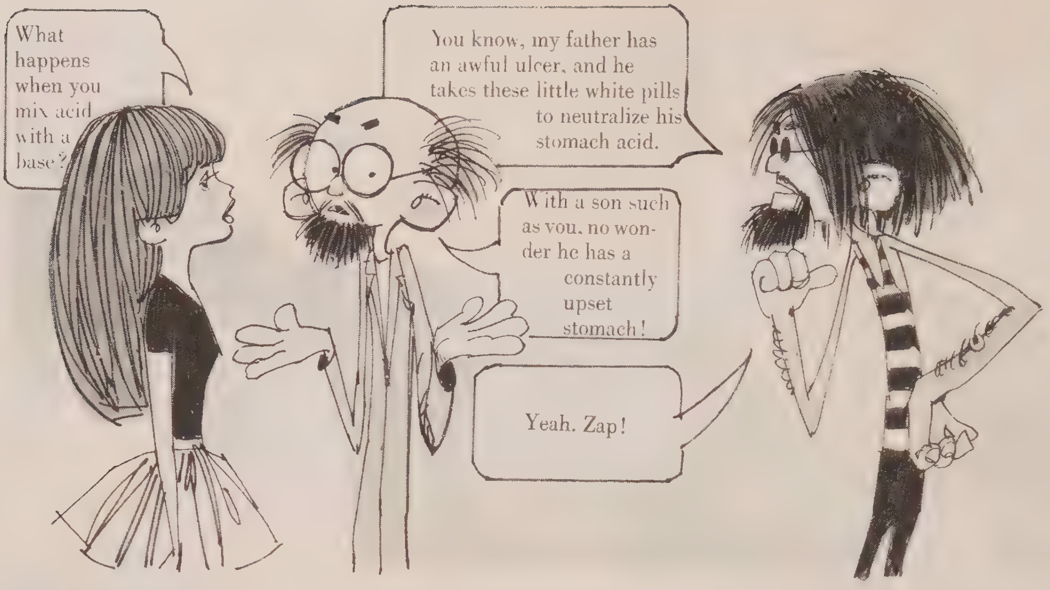
1. Complete the word equations for the reactions you caused in Part II:

(a) sodium hydroxide + iron (III) nitrate → ______________________ + ______________________

(b) sodium hydroxide + cobalt (II) nitrate → ______________________ + ______________________

(c) sodium hydroxide + zinc nitrate → ______________________ + ______________________

2. Why isn't sodium or potassium stored in water instead of in kerosene? ______________________________


What happens when you mix acid with a base?
You know, my father has an awful ulcer, and he takes these little white pills to neutralize his stomach acid.
With a son such as you, no wonder he has a constantly upset stomach!
Yeah. Zap!

Neutralization

What happens when an acid and a base are mixed? Find out in this experiment.

Experiment 8.3

Equipment and Supplies per Lab Station

beaker—250 ml; dropper, long stem; evaporating dish
graduated cylinder—10 ml; pipet-(5 or 10 ml) to measure 1 ml
watch glass; asbestos wire gauze; ring stand with ring
litmus paper, red and blue; water, distilled or deionized
Solutions of: hydrochloric acid; phenolphthalein in dropping bottle; sodium hydroxide

What to do

1. Transfer 5 ml of 1M sodium hydroxide from a pipet into a clean evaporating dish. Add a drop of phenolphthalein. What is the color of the indicator in the base?

   ____
2. In a 10-ml graduated cylinder measure 6.0 ml of 1M hydrochloric acid.
3. Slowly add 2 ml of the acid solution to the dish containing the sodium hydroxide solution. Feel the underside of the dish. Inspect the color of the indicator? What do you observe, if anything? ______________________________

   ____
4. Keep on stirring. Use a long stem dropper to add more acid, a drop at a time, until the color *just* disappears. See Fig. 34. How much of the 1M acid did you use? ______________________________
5. Pour about 2 ml of the resulting solution onto a watch glass. Evaporate this liquid to dryness. (See HOW TO EVAPORATE A LIQUID in Chapter 4.)

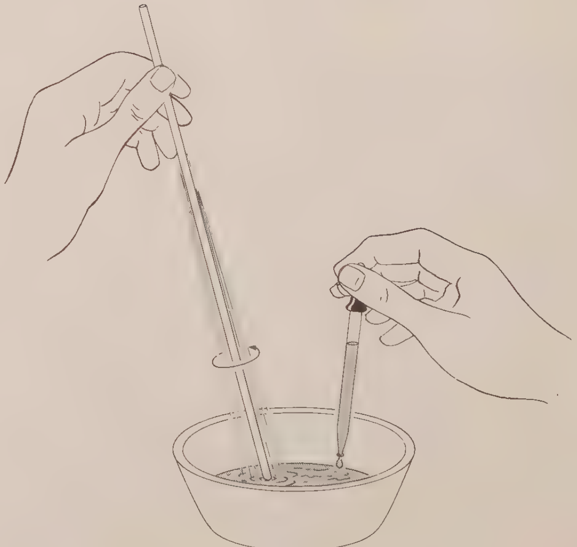


Fig. 34.

6. Cool the watch glass. Dissolve the solid on the glass in 2 ml of distilled water.
7. Test the dissolved solid with both red and blue litmus. Write down any observations. __________

____

Is this solid an acid? ____________ a base?

____

Questions to answer

1. The reaction you have observed in this experiment is called neutralization. In proper amounts, an acid makes a base lose its properties or a base makes an acid lose its properties. When an acid and a base neutralize each other, the products formed are a salt and water.

How would you define a salt? ______________

____

2. The name of a salt is composed of two parts. The first part is like the beginning of the base's name. The last part of the salt's name comes from acid used. Complete the table below. The Reference Table in the APPENDIX shows you the acid ending in a salt's name.

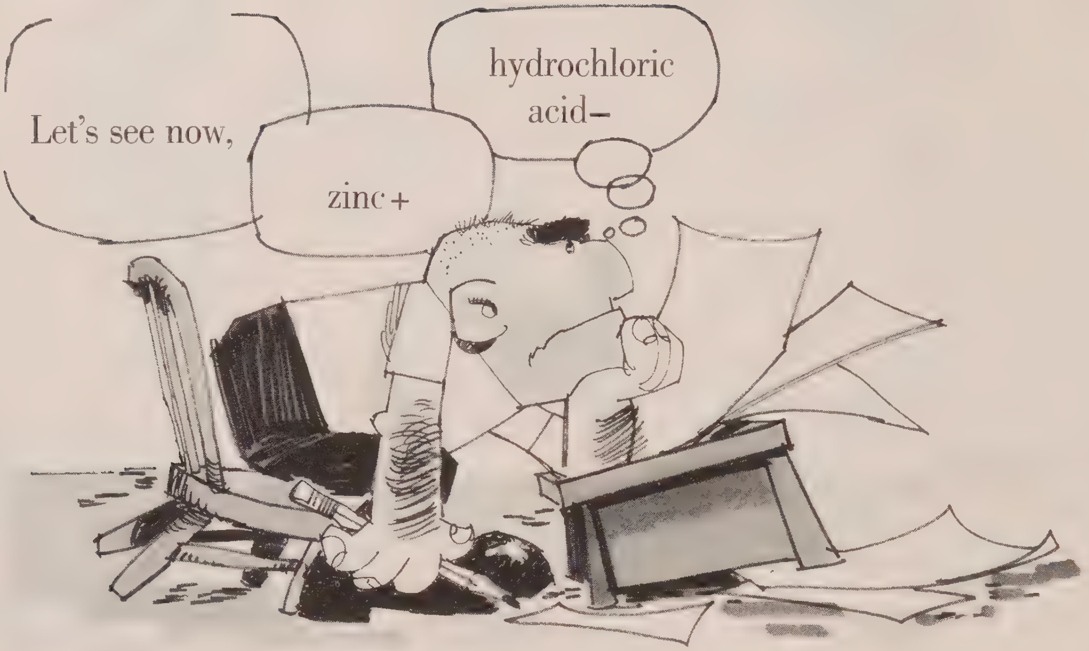
NAME OF SALT	MADE BY NEUTRALIZATION OF	
	Base	Acid
1. sodium chloride		
2.	sodium hydroxide	sulfuric acid
3. potassium chloride		
4. potassium nitrate		
5.	calcium hydroxide	nitric acid

Extra for experts

Complete the word equation for the neutralization reactions:

1. sodium hydroxide + nitric acid → ______________ + ______________
2. calcium hydroxide + acetic acid → ______________ + ______________
3. barium hydroxide + hydrochloric acid → ______________ + ______________
4. magnesium hydroxide + sulfuric acid → ______________ + ______________
5. potassium hydroxide + phosphoric acid → ______________ + ______________

Exercise



How are you doing?

Circle the best statement or answer to each question. Write your answer in answer blank.

1. Which would you expect to turn red litmus blue?
   (a) hydrochloric acid
   (b) sodium hydroxide
   (c) table salt
   (d) distilled water

1. ____________

2. Which color change occurs when phenolphthalein is put into a base?
   (a) red to blue
   (b) orange to yellow
   (c) orange to red
   (d) colorless to red

2. ____________

3. Which would you react to produce sodium chloride?
   (a) sodium and water
   (b) sodium hydroxide and hydrochloric acid
   (c) sodium hydroxide and water
   (d) hydrochloric acid and water

3. ____________

4. What would you add to sodium bicarbonate (baking soda) to produce a gas?
   (a) sodium hydroxide
   (b) water
   (c) hydrochloric acid
   (d) sodium chloride

4. ____________

5. What two substances would you react to obtain sodium sulfate?
   (a) sodium and water

   (b) sodium hydroxide and hydrochloric acid

   (c) sodium hydroxide and sulfuric acid

   (d) sulfuric acid and ammonia water

5. ________

6. Which pH indicates the strongest base?
   (a) 1

   (b) 9

   (c) 7

   (d) 13

6. ________

7. Which is an *incorrect* statement about bases?
   (a) Most bases are of low solubility in water.

   (b) Most bases are made by neutralization.

   (c) Bases can destroy skin tissue.

   (d) Bases react with hydrochloric acid.

7. ________

8. Which would you classify as a salt?
   (a) sodium hydroxide

   (b) sodium chloride

   (c) household ammonia

   (d) hydrochloric acid

8. ________

9. Which combination will produce water as one of the products?
   (a) iron and silver nitrate

   (b) potassium hydroxide and nitric acid

   (c) copper and sulfur

   (d) copper (II) chloride and aluminum nitrate

9. ________

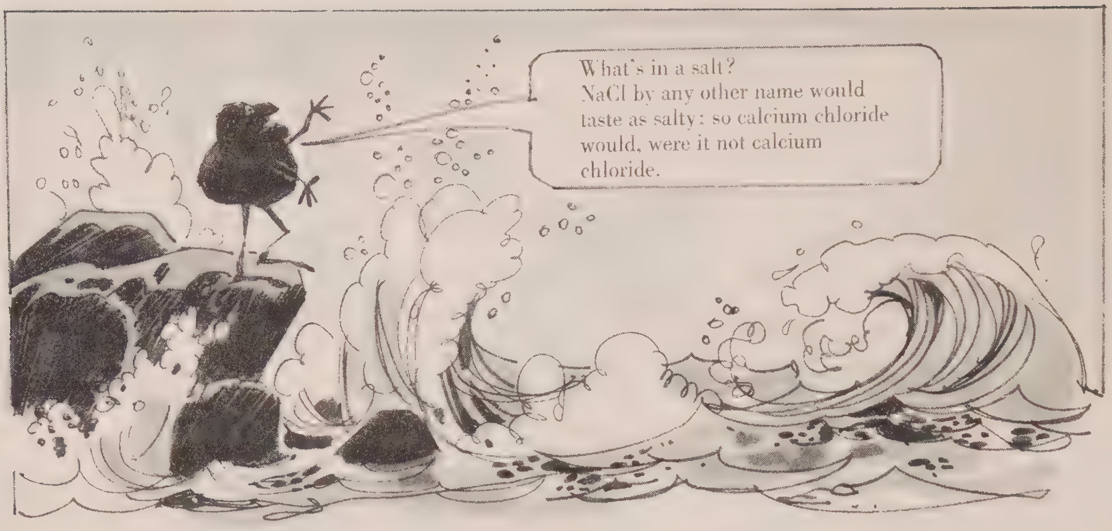
10. The indicator which turns from blue to red in acids is called
   (a) litmus

   (b) phenolphthalein

   (c) universal indicator

   (d) bromphenyl blue

10. ________


What's in a salt?
NaCl by any other name would taste as salty: so calcium chloride would, were it not calcium chloride.

Synthesis of a Salt

Some salts are found in the earth while others have to be made from other substances. In this experiment, you will use several reactions to obtain the salt, calcium chloride, as a final product.

Experiment 8.4

Equipment and Supplies per Lab Station

beaker—250 ml; beaker—100, 150, or 250 ml
dropper; evaporating dish; graduated cylinder—25 ml
stirring rod; (2) test tubes; asbestos wire gauze
burner; forceps; ring stand with ring; test tube rack
calcium metal; water, distilled or deionized
Solutions of: hydrochloric acid
phenolphthalein indicator in dropping bottle

What to do

1. Pour 25 ml of water into a small beaker.
2. Use forceps to add a small (0.5—1 cm long) piece of calcium metal to the water. See Fig. 35. What evidence is there that a reaction is taking place? ____________________

   ____
3. After the reaction is completed, pour about 1 ml of the beaker's content into a test tube. Add a drop of phenolphthalein indicator to the test tube. Record your observations. ____________

   ____

   Does the phenolphthalein give either an acid or base result? ____________________

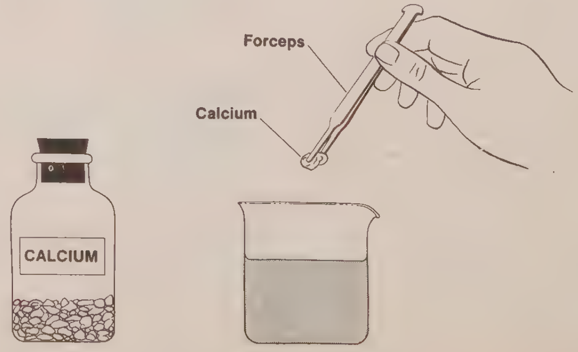


Fig. 35.

4. Add just enough hydrochloric acid, with a dropper to make the liquid in the beaker clear. Is there any evidence that a reaction is taking place? Explain. ______________________

____

____

5. Add 5 more drops of hydrochloric acid to the beaker.
6. Pour about 1 ml of the liquid in the beaker into a second test tube. Add a drop of phenolphthalein. Record your observations.

____

____

What does the phenolphthalein show to be present in the beaker now?________________

7. Put 10 ml of the liquid in the beaker into an evaporating dish. Evaporate the liquid to obtain some of the salt you just made. Describe the appearance of the salt.___________________

____

____

Questions to answer

1. Write the word equation for the reaction that occurs in step 2. (It is a single replacement reaction in which water acts as if it is hydrogen hydroxide (HOH). Underline the product from this reaction. The product is used in step 4.

____

____

2. Write the word equation for the double exchange reaction that occurs in step 4._______________

____

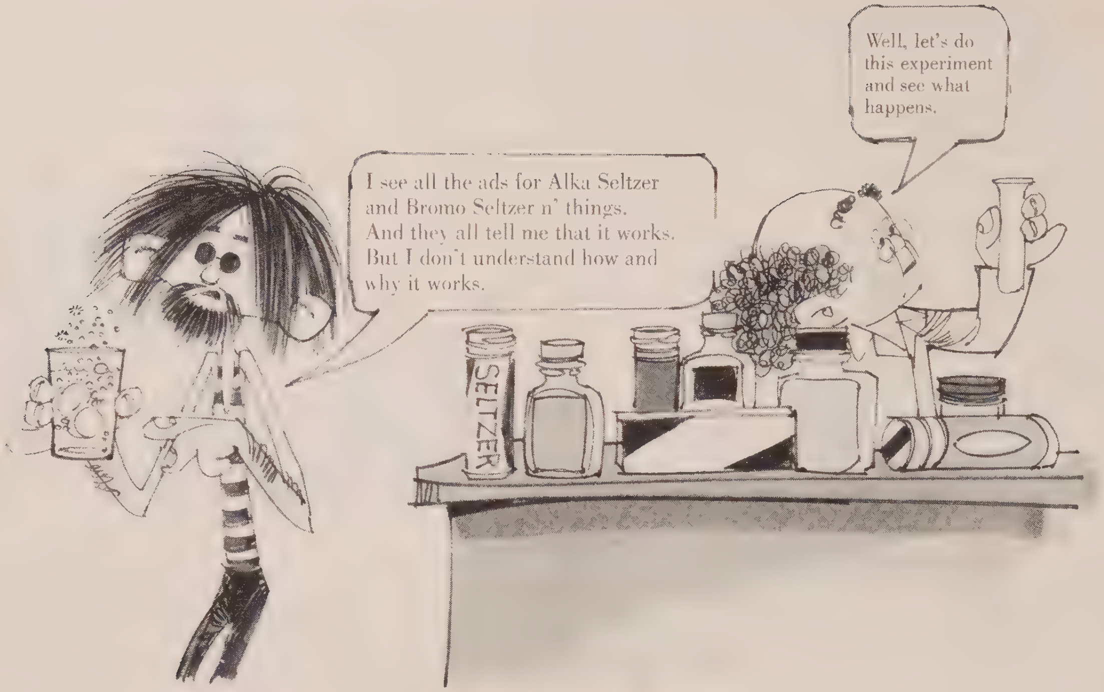
____

Extra for experts

What are two other ways by which calcium chloride might be made? ______________________________

____

____


I see all the ads for Alka Seltzer and Bromo Seltzer n' things. And they all tell me that it works. But I don't understand how and why it works.
Well, let's do this experiment and see what happens.
SELTZER

The Anti-Burps

TV commercials lead us to believe that some products can be used to overcome gastric acidity. Are they telling the truth?

Experiment 8.5

Equipment and Supplies per Lab Station

beaker 100, 150, or 250-ml
dropper
hydrochloric acid approximately 1*M*
methyl orange indicator
antacid 1
mortar and pestle
antacid 2
stirring rod
antacid 3
water, distilled or deionized

What to do

1. Pour 10 ml of hydrochloric acid into a small beaker.
2. Put 2 or 3 drops of methyl orange indicator into the acid. Observe the color of the indicator in an acid. Write your observation in the chart below.
3. With a dropper add a few drops of antacid 1 to the acid. Stir the liquid for a minute or so. Record any changes you see.
4. Continue to add drops of antacid 1 and to stir until you do not see any more changes taking place. Write any additional observations on the chart.

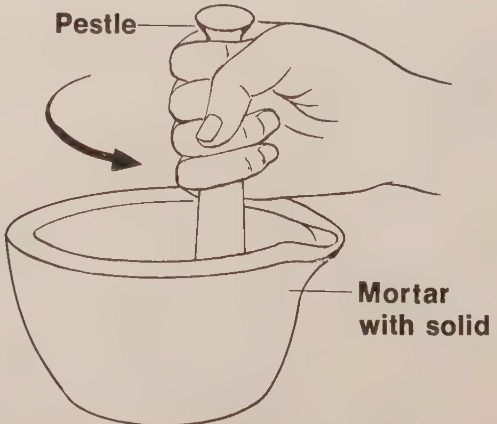


Fig. 36. Grind the solid in a mortar to make a fine powder.

Do you think that antacid 1

counteracted the effects of the acid? ___________

Why? ________________________________

____

5. Your teacher will show you how to mix tablets with water so they can be used in a dropper. See Figs. 36 and 37.
6. Repeat steps 1–4 using antacid 2 instead of antacid 1. Write all your observations in the chart.
7. Repeat steps 1–4 using antacid 3 instead of antacid 1. Record your observations in the chart.

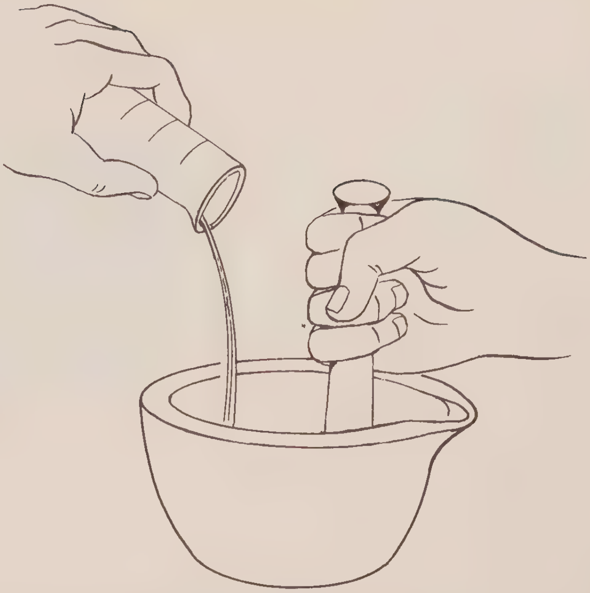


Fig. 37. Gradually add 10 ml of water as you continue to grind the powder.

DATA CHART

Color of indicator in acid	Substance added to acid	Any change in color of indicator	Any other observations

Questions to answer

1. Do you think that the acid was counteracted by each of the substances used? ______________

   Explain. ________________________________

   ____

   ____

   ____

2. Are antacids 1, 2, and 3 the same substance?

   ____

   What are your reasons for thinking this?

   ____

   ____

Acid-Like and Base-Like Substances in the Home

There are many acid-like and base-like substances found in a home. Which of those given you have acid-like or base-like properties?

Equipment and Supplies per Lab Station

dropper
litmus, both red and blue
spoon
stirring rod
4 or more substances to be tested
universal indicator papers
watch glass
water, distilled or deionized

What to do

1. Select one of the liquid substances. Then with a clean stirring rod transfer a drop of the liquid to each color of litmus papers. Write your observations in the data column.
2. Transfer a drop of the liquid selected to universal indicator paper. Compare the color with the color-pH reference chart and determine the pH of the substance.
3. Clean and dry the stirring rod before using it in another solution.
4. Repeat steps 1–3 until you have tested the number of liquids designated by your teacher. Add all of your observations to the data chart.
5. Select one of the solids. Put about 1/8 spoonful of the solid on a watch glass. Add a few drops of water and stir the mixture.
6. Repeat steps 1–3 using the solution you just made instead of one of the liquids. Record your observations in the data chart.
7. Continue to test the other solids as you did in steps 1–3. Complete the Data Chart with your observations.

DATA CHART

Substance	Effect on		Universal Indicator	
	red litmus	blue litmus	color	pH

Questions to answer

1. On the basis of your observations, which substance(s) is (are) acid-like? ______________________
____
____
____
2. Which substance(s) is (are) base-like? ______________
____
3. Which substance(s) might be classified as "neutral"? ______________________________
Why? ______________________________________
4. Which substance, of those tested, was the strongest base? ______________ the strongest acid? ______________ Explain why you chose these substances ______________________________
____

Laboratory Experiments in

ACTION CHEMISTRY

UNIT IV

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Chapter 9

ELECTRIC FORCES

Electricity is a form of energy. You can see what electricity does. You cannot see electricity itself.

In this demonstration you will discover answers to the following questions about electric forces:

How many kinds of electricity are known?

Are all objects and particles of matter electrically-charged?

If not, how can you tell uncharged objects from charged ones?

How do electrically-charged objects affect each other?

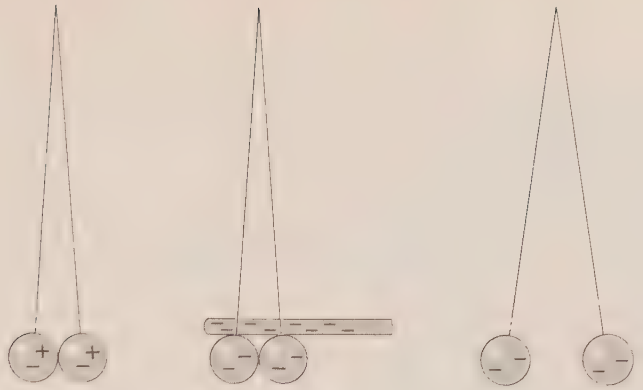
Demonstration 9-A

Equipment and Supplies per Lab Station

hard rubber rod; piece of fur; glass rod
piece of silk; (2) ring stands; (2) clamps
(2) pith balls attached to nonconducting threads

What to do

1. Try to pick up bits of paper with a hard rubber rod. Does the paper stick to the rod? ______________________
2. Rub the hard rubber rod vigorously with fur. Now try picking up the bits of paper. What happens? ______________________
3. Repeat steps 1 and 2 with a glass rod. Rub the rod with silk. What happens
   (a) with the rod *before* being rubbed? ______________________
   (b) with the rod *after* being rubbed? ______________________
4. Suspend two pith balls as illustrated in Fig. 38(a).
5. Touch both balls with the hard rubber rod which has been rubbed with fur. What happens? See Fig. 38(b). ______________________



(a) Suspended pith balls

(b) Rubber rod touching both pith balls.

(c) Pith balls after being touched with rubber rod.

Fig. 38.

6. Touch each ball with your finger. What happens? ______________________________

____

7. Repeat step 5 with the glass rod rubbed with silk. Again, what happens? ______________________________

____

8. Again touch each ball with your finger. Why? ______________________________

9. Now touch one ball with the hard rubber rod rubbed with fur. Touch the other ball with the glass rod rubbed with silk. How do the balls behave this time? ______________________________

____

Questions to answer

1. How many kinds of electricity have you discovered? ______________ Explain. ______________

____

____

2. Are all objects electrically charged? ______________ If not, how can you tell uncharged objects from charged ones? ______________________________

____

3. How do electrically charged objects affect each other:

   (a) Those with like charges? ______________

   (b) Those with unlike charges? ______________

Extra for experts

1. Negative and positive electric charges

   (a) *The charge produced on a hard rubber rod by rubbing it with fur or wool has been called NEGATIVE.*

   (b) *The charge produced on a glass rod by rubbing it with silk has been called POSITIVE.*

2. See if you can find out how scientists explain the charging of the rods. ______________________________

____

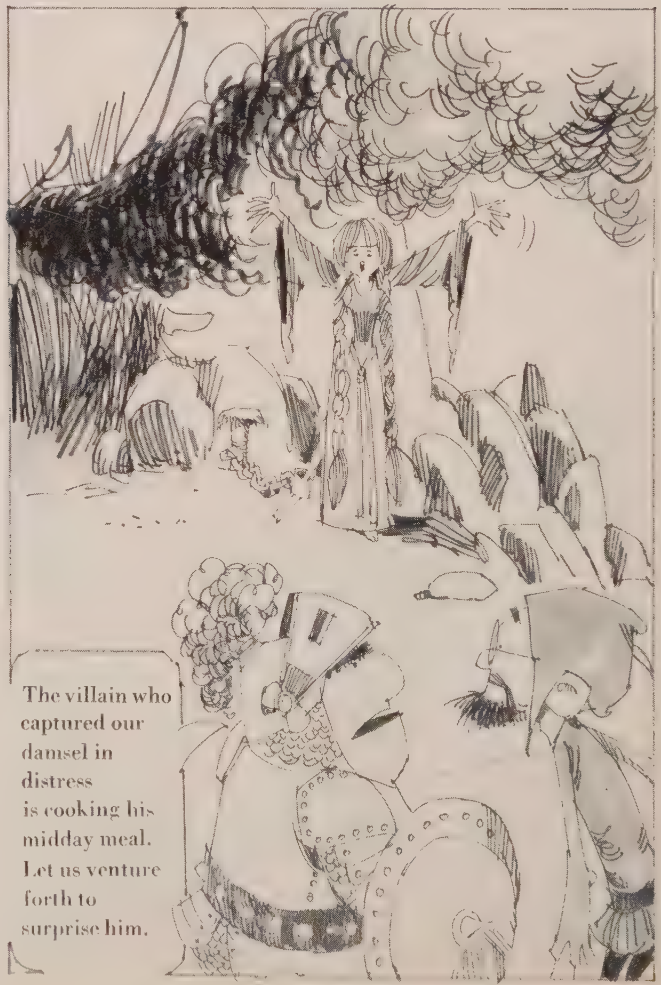
MYSTERY BOXES

You can never really see atoms. Why then believe in atoms? Scientists have used *indirect* methods to "see" atoms. They have learned a great deal about how atoms behave and how they are made.

In this experiment you will use an indirect method. Learn all you can about objects you cannot see. Use your evidence to describe the objects, without actually seeing the objects.

NOTE: (1) Do not try to guess what the object actually is.
(2) It is against the rules to open or cut into the box.

Experiment 9.1


The villain who captured our damsel in distress is cooking his midday meal. Let us venture forth to surprise him.

Equipment and Supplies per Lab Station
Mystery boxes

What to do

1. Obtain a numbered, sealed "mystery" box. Record the number in the Data Chart.
2. Shake carefully, tilt, and otherwise move the box. Listen for sounds. Notice the "feel" of the box in different positions. Notice whether there seems to be one or more objects in the box. Observe how the object(s) inside seem(s) to move.
   (a) Record your observations in the space provided.
   (b) Describe the object or objects: heavy or light, rounded or irregular in shape, metal, glass, or plastic; any other ideas you may have.
3. Trade boxes with your classmates. Make observations on as many boxes as time permits.

Questions to answer

1. Compare your observations on individual boxes with those of your classmates. Do your observations agree? ______________________
____
2. Check with your teacher on how close your observations were to what might be expected.
____
____
3. How much information can you get by an indirect method? ______________________
____

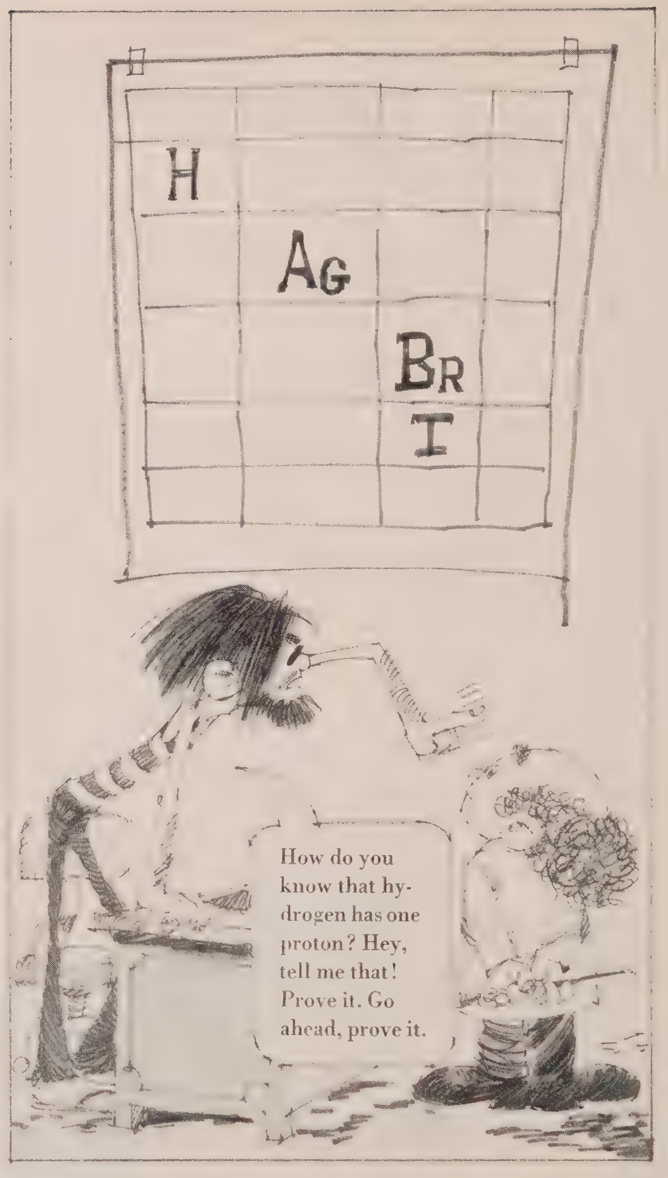
(Clara Aich and George Senty)

DATA CHART

"MYSTERY BOXES"

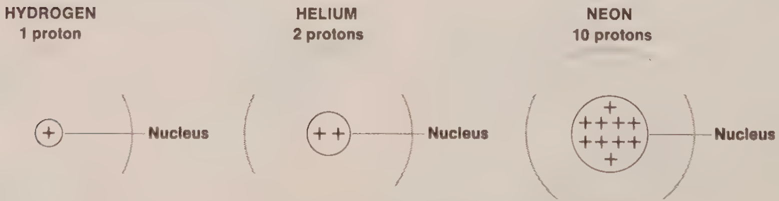
Box Number	How Many Objects?	Shaking	Tilting	Other Observations	Description

Making a Periodic Table


H
Ag
Br
I
How do you know that hydrogen has one proton? Hey, tell me that! Prove it. Go ahead, prove it.

In the Periodic Table of the Elements, a great deal of information about the elements is organized. The elements are arranged in order of their increasing atomic numbers, from Atomic Number 1 to Atomic Number 105.

1. *ATOMIC NUMBER is the number of protons in the atom.* Each proton has one positive electric charge.
2. An electrically neutral atom contains one negative electric charge, called an *electron*, for each of its protons. Atoms also contain neutral particles called neutrons. You will learn more about electrons and neutrons in later Exercises.
3. *The NUCLEUS of an atom is an area in the atom into which protons and neutrons are tightly packed.* See Fig. 39. Electrons move about in the space in the atom around the nucleus. When elements are arranged in order of increasing atomic numbers, an element is found every so often which has properties similar to those of an element which came before it. Chemists put elements with similar properties in the same vertical column in the table. These groups are called FAMILIES. Look at the Noble Gas Family. See Fig. 40. You will notice other families when you are making your Periodic Table. One of these is the Alkali Metal Family. See Fig. 41.



Protons each carry one unit of positive electricity and are represented by a "+" sign.

Fig. 39.

Noble gases VIII-A	
2 **He**	He is used in air balloons and in diving equipment.
10 **Ne**	Ne is used in neon signs.
18 **Ar**	Ar is used to fill light bulbs.
36 **Kr**	Kr is used in high-speed electronic flash bulbs.
54 **Xe**	Xe is used in high-speed electronic flash bulbs.

Fig. 40. The Noble Gas Family — The Unsociables
Properties of this family: colorless, odorless gases which do not react readily with other elements.

Alkali Metals I-A	
3 **Li**	Li is used in ceramics, welding, and drugs
11 **Na**	Sodium compounds include table salt, sodium chloride
19 **K**	Potassium compounds are used in fertilizer
37 **Rb**	Rb is used in photoelectric cells
55 **Cs**	Cs is used in photoelectric cells

Fig. 41. The Alkali Metal Family — The Very Sociables
Properties of this family: soft white metals, which react so readily with other elements that only compounds of these metals are found in nature. No free metals in this family are found in nature, whereas copper and silver are found uncombined.

What to do

1. Obtain four and one-half feet of one-inch wide adding machine tape. Mark this off into 54 one-inch squares.
2. Number the squares — 1, 2, 3, 4, etc. up to 54. Place the numbers in the upper left-hand corners of the squares.
3. Cut off squares 1 and 2. Put them aside.
4. Now cut the tape between squares 10 and 11.
5. Cut the tape between squares 18 and 19.
6. Cut the tape between squares 36 and 37. Leave the squares from 37 to 54 in one piece.
7. Tape two sheets of loose-leaf notebook paper ($8^1/_2$ in × 11 in) together along the bottom ends. See Fig. 42.

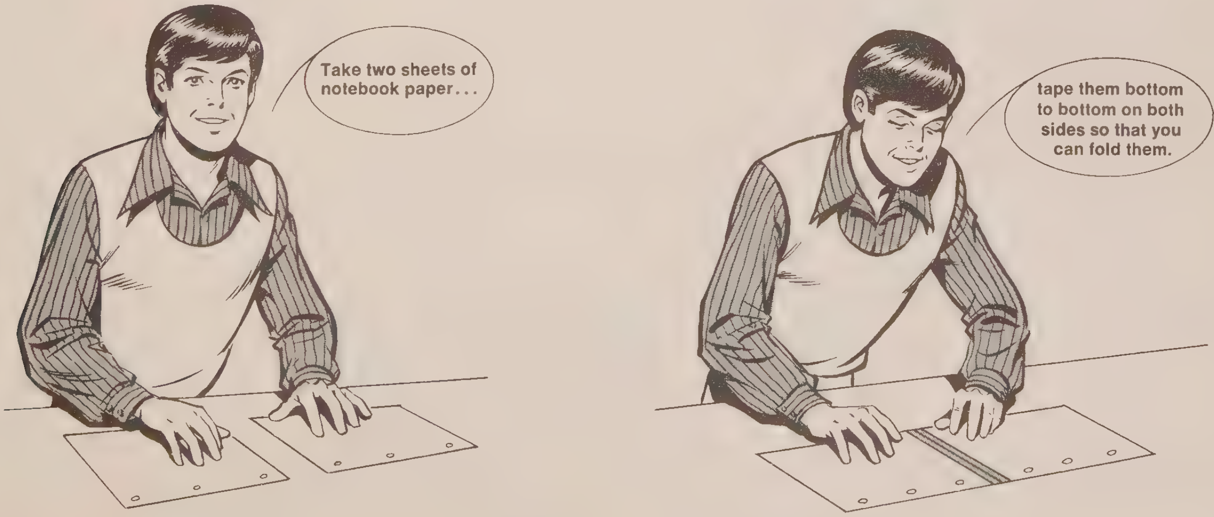

Take two sheets of notebook paper . . .
tape them bottom to bottom on both sides so that you can fold them.

Fig. 42.

8. Tape the row of squares 37 to 54 along the lower edge of the sheets.

9. Line up the squares of the Noble Gases above 54. See Fig. 40. Tape these in place.

10. Now cut the row of squares 11 to 18 between squares 12 and 13.

11. Cut the row of squares 3 to 10 between squares 4 and 5.

12. Line up the squares of the Alkali Metals above 37. See Fig. 41.

13. Cut squares 1 and 2 apart. Place square 1 above square 3. Place square 2 above square 10.

14. Tape all of your squares to the sheets of paper.

15. Label your pages across the top: PERIODIC TABLE OF THE ELEMENTS.

16. Use Table 4 in the APPENDIX to fill in the symbol for each element. Write the correct symbol in the center of each square.

17. Periods are the horizontal rows in the Periodic Table. Number them from the top down, 1, 2, 3, 4, 5.

18. Mark "A" families of elements in the vertical columns as follows:

I-A	II-A	III-A	IV-A	V-A	VI-A	VII-A	Noble Gases
H	Be	B	C	N	O	F	He
etc.	etc.	etc.	etc.	etc.	etc.	etc.	etc.

Refer to Table 4 in the APPENDIX.

19. Fold your Periodic Table and place it in your notebook to complete later.

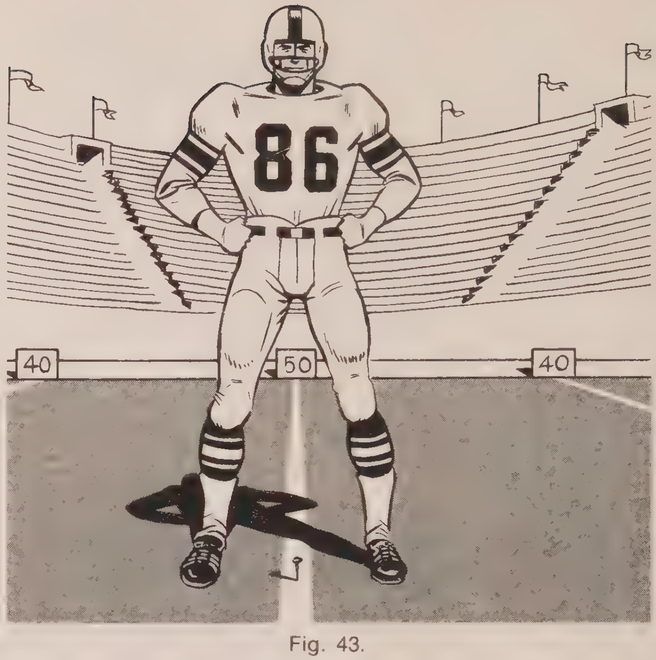


Fig. 43.

Questions to answer

1. What was the fewest number of elements found in any one period? ______________ the largest number? ______________________________

2. Which periods have the same number of elements? ______________________________

____

3. The "B" families of elements are known as the *transition elements.* Between what "A" families are they found in the Periodic Table? ________

____

Adding to The Periodic Table

Electrons in the atom are whizzing about at great speed. They must fill the space around the nucleus of the atom. This space is like a large football stadium. In proportion, the nucleus is like the head of a small pin. See Fig. 43.

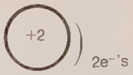
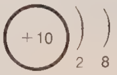
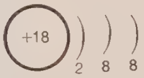
The electrons travel in pathways called ORBITALS. These orbitals are arranged in areas around the nucleus called *ENERGY LEVELS.* It is believed that there are at least seven energy levels. Energy level *one* is close to the nucleus. It has a low amount of energy. Each energy level farther away from the nucleus with a higher *number* has higher energy.

Can electrons change energy levels? Yes, if they can take on extra energy or release energy. You will learn more about this later.

How do you map energy levels? One way is to represent them by means of arcs of a circle placed at increasing distances from the nucleus. The nucleus of the atom is represented by its positive charge. Study the table which follows, showing energy levels in three atoms.

How many electrons can move in any one energy level? A definite maximum number is allowed. Energy levels start filling close to the nucleus. When one is filled, a new one must be started. The maximum number of electrons for each energy level will be found in the Noble Gases. The energy levels of the first three Noble Gases may be pictured as follows.

ELECTRONS AND ENERGY LEVELS

Noble Gas	Atomic Number	Electrons found in			*Arc Representation
		Energy Level 1	Energy Level 2	Energy Level 3	
Helium	2	2			
Neon	10	2	8		
Argon	18	2	8	8	

*Note: The total number of electrons equals the number of protons.

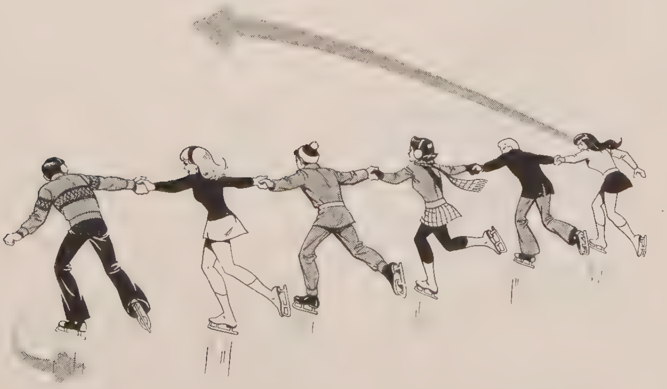


Fig. 44. The skater farthest from the center requires the most energy to skate a circle.

What to do

PART I REPRESENTING ENERGY LEVELS

1. Use the energy levels of the Noble Gases as a guide. Draw the arc representation for each of the *first twenty elements only* across the bottom of the square in your Periodic Table.
2. Notice that
   (a) lithium and all of the elements of Period 2 have one more energy level than hydrogen and helium.
   (b) the elements of Period 3 have one more energy level than neon.
   (c) the elements of Period 4 have one more energy level than argon.

PART II LOCATING METALS, NONMETALS, AND METALLOIDS

1. Draw a heavy line down the left side and across the bottom of the square for boron, B. Now do the same for the squares for silicon, Si; arsenic, As; and tellurium, Te. You will have a jigged line. Compare with Table 4 in the APPENDIX.
2. All elements to the left and below the line are metals. Those to the right and above the line are nonmetals. Those along the line are metalloids. *METALLOIDS are elements having certain properties characteristic of metals, but are generally classed as nonmetals.*

So, that's what a hydrogen atom looks like!

Questions to answer

1. How many electrons are there in the highest (outermost from the nucleus) energy level in each of the Group A families of elements?

Group I-A ____________ Group V-A ____________

Group II-A ____________ Group VI-A ____________

Group III-A ____________ Group VII-A ____________

Group IV-A ____________ Noble Gases ____________

3. How does the number of metals compare with the number of nonmetals? ____________________

____

4. How can the Periodic Table help you in understanding and remembering the properties of elements? ____________________

____

COLORED FLAMES

What makes the colors in a driftwood fire at the beach? What is used to color the flame of a wood fire in the fireplace?

How does a chemist explain the bright colors you see in the driftwood fire? In this experiment?

In the Exercise, *Adding to the Periodic Table*, you learned that when an electron in an atom gains enough energy it can jump from a lower energy level to a higher one. Then, as it loses this extra energy, it falls back to the lower energy level. The extra energy lost appears as colored light. Each color represents a definite quantity of energy in the form of light. The electrons of different atoms require different amounts of energy to jump to higher energy levels. They give off this energy as different colors of light as they fall back to the lower levels of energy.

Experiment 9.2

Equipment and Supplies per Lab Station

set of eight bottles, heavy glass, wide-mouthed, each fitted with a cork through which is inserted a glass rod containing a platinum wire—see Fig. 43.
cobalt glass squares; laboratory burner; unknown solution
Solutions of: sodium nitrate; lithium nitrate;
strontium nitrate; calcium nitrate;
barium nitrate; potassium nitrate; and
potassium nitrate solution + sodium nitrate solution

What to do

1. Obtain a set of solutions containing the nitrates of the metals listed in the Data Chart. Each solution will be supplied as shown in Fig. 45.
2. Set up a laboratory burner with a flame with a blue inner cone and no yellow.
3. Dip the tip of the wire in the solution in the bottle. Place just the tip of the wire in the light blue burner flame. Record your observations in the Data Chart.
4. Repeat the flame test with each of the samples in the set. Record your observations on each sample.
5. Do the flame test again with the sodium nitrate solution. Look at the colored flame through two thicknesses of cobalt glass. Record your observations.

Why does the sun change color when it's near the horizon? The flames of the sun don't change color do they?

6. Look at the flame with potassium nitrate solution through two thicknesses of cobalt glass. Record your observations.
7. Do the flame test with the mixture of potassium nitrate and sodium nitrate solutions. Record your observations.

Now look at the flame of the mixture through the two thicknesses of cobalt glass. Record this additional observation.

8. Ask your teacher for an unknown. Decide what metal is in your unknown. Record your observations supporting your decision.

The unknown solution contains ______________

Right or wrong? ______________

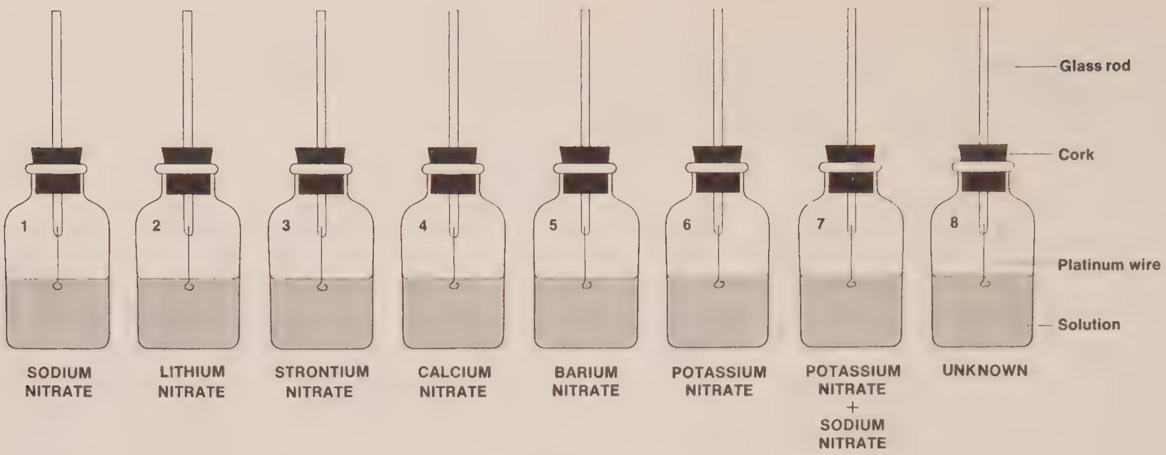


Fig. 45.

DATA CHART

FLAME TESTS	
Metal in the Solution Used	Observations
Sodium	
Lithium	
Strontium	
Calcium	
Barium	
Potassium	
Sodium, with cobalt glasses	
Potassium, with cobalt glasses	
Sodium and potassium with cobalt glasses	
Unknown	
Unknown, with cobalt glasses	

Questions to answer

1. How does the cobalt glass make it possible to test for potassium in a solution which also contains sodium? ______________________

2. What difficulty may be found in using the flame test? ______________________

3. How does the chemist use the flame test? ______________________

Extra for experts

What information about the electrons in the metal particles in the solution does the color of the flame give the chemist? ______________________

USING A SPECTROSCOPE

Can you tell what is in "brand X"?

Experiment 9.3

Equipment and Supplies per Lab Station
evaporating dish; (2) asbestos pads; laboratory burner
spectroscope
ethyl alcohol; sodium chloride; sodium bicarbonate
Salts of: potassium; barium; strontium; calcium; copper
unknown materials

What to do

1. Put 2 or 3 ml of alcohol in an evaporating dish. Set the dish on an asbestos pad.
2. Ignite the alcohol. Observe the color of the flame. Write your observation in the Data Chart on page *4*/14.
3. Look through the spectroscope at the burning alcohol. Write any additional observations in the Data Chart.
4. Smother the first by putting another asbestos square over the top of the evaporating dish.
5. Put about ¼ spoonful of sodium chloride in a clean evaporating dish.
6. Add 2-3 ml of alcohol to the sodium chloride.
7. Ignite the alcohol. Write in the Data Chart the color of flame you observe.
8. Look through the spectroscope at the burning alcohol—sodium chloride mixture. Note down in the Data Chart what you see.
9. Repeat steps 5-8 but use sodium bicarbonate instead of sodium chloride. Add your observations to the Data Chart.
   Do you think you would get the same colored flame and spectroscopic results if you use

   some other sodium compound? ____________

   ____

   Why do you think this? ____________

   ____

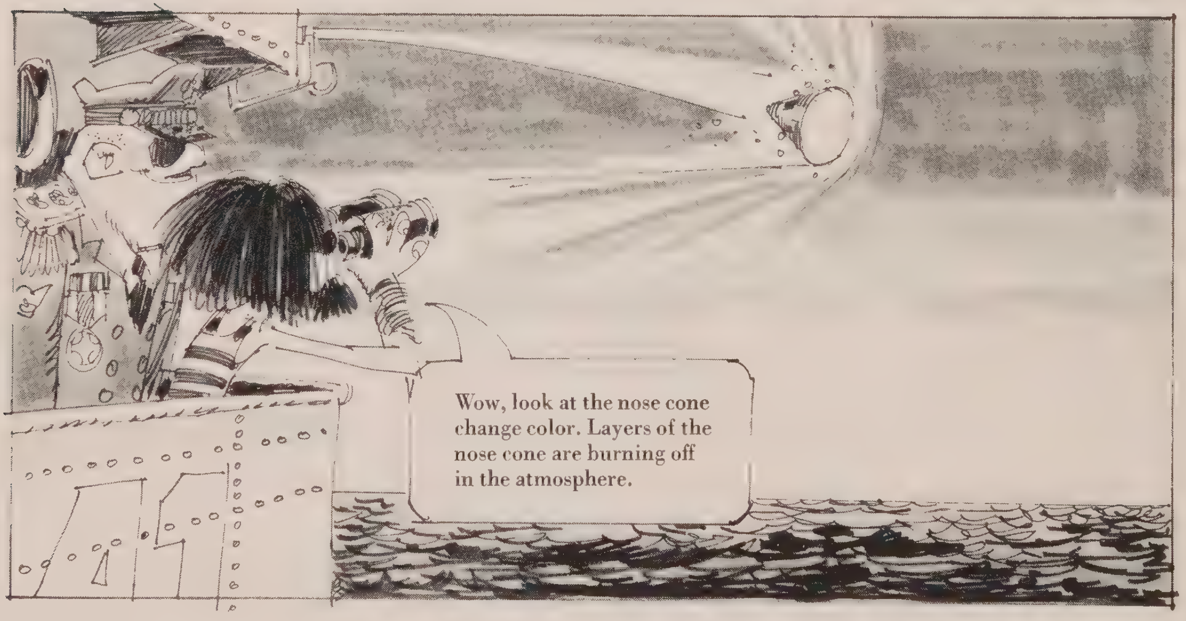
   ____

10. Repeat steps 5-8 several times but use a different salt with the alcohol each time. Write all observations in the Data Chart.
Do you think a spectroscope can be used to identify an unknown substance? ______________
____

11. Repeat steps 5-8 with an unknown solid. Record your observations.
What do you think the solid is? ______________
Why do you think this? ______________
____

DATA CHART

WHAT YOU HAVE SEEN		
Substance Alcohol	Color of flame	Seen through spectroscope
Alcohol and sodium chloride		
Alcohol and sodium bicarbonate		


Wow, look at the nose cone change color. Layers of the nose cone are burning off in the atmosphere.

Questions to answer

1. Why do you think the burning alcohol was used with the salt? ______________
____
Do you think a burner could have been used in place of the alcohol? ______________ Why do you or do you not think so? ______________
____

2. Without looking into the spectroscope, how could you find out what spectrum was being produced by a substance? ______________
____

3. When do you think a chemist would use a spectroscope to identify a substance? ______________
____
____

Chapter 10

THE GROUP

The outermost energy level in an atom may be called the VALENCE ENERGY LEVEL. Electrons found in the valence energy level are called VALENCE ELECTRONS.

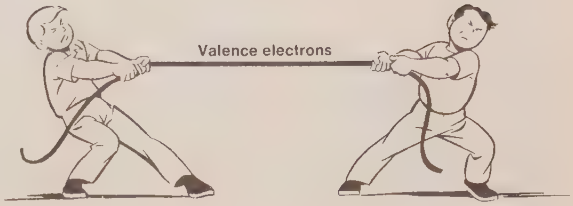
The valence energy level is filled when it contains eight electrons. Most atoms do not have completely filled valence energy levels. One way that atoms may obtain completely filled valence energy levels is by sharing electrons with one or more other atoms.

Atoms are held together by attractive forces called CHEMICAL BONDS. One kind of chemical bond is called the covalent bond. *A pair of shared electrons produces a COVALENT BOND holding two atoms together.*

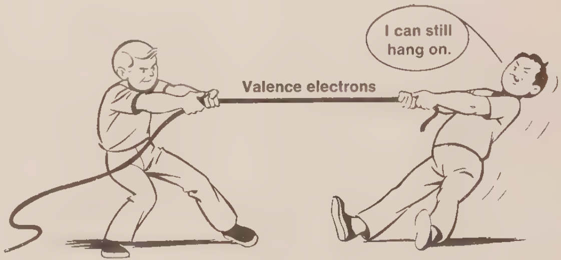
The sharing of electrons is possible because the electrically positive nucleus of each atom is attracting not only its own electrons but also the electrons of the other atom. If the attracting force of the two nuclei is equal, the electrons are shared equally. This is said to be *non-polar bonding.* However, one nucleus may have a stronger pull on the shared electrons than the other nucleus. The electrons will be shared unequally. This bonding is called *polar.* If the pull is strong enough, there may not be any sharing of electrons, but rather the forming of still another type of bonding which you will study in Experiment 10.2.

Atoms of elements found near each other in the Periodic Table tend to form covalent bonds. Atoms of elements appearing farther apart in the Periodic Table tend to form another type of chemical bond.

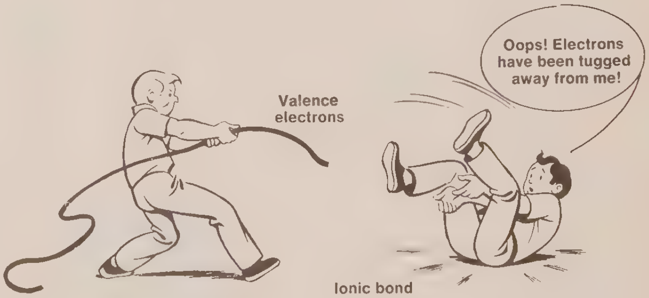
In this experiment you will make models of some compounds which contain covalent bonds.



Nonpolar covalent bond



Polar covalent bond


Oops! Electrons have been tugged away from me!
Valence electrons

Ionic bond

Fig. 46.

Experiment 10.1

Equipment and Supplies per Lab Station

1 ball and stick molecular model kit

What to do

1. Obtain a ball and stick model kit. Inspect its contents. Each ball represents an atom. Different colored balls represent different kinds of atoms. Each hole in a ball represents a valence electron. The hole is a place at which the atom may form a covalent bond with another atom.
2. Select a ball having four holes, one having two holes, and one having only one hole.
3. Complete the chart below to identify the element each ball represents.

MODEL CHART

Element	Symbol	No. of valence electrons	No. of holes in the ball	Color of the ball
Carbon	C	4		
Hydrogen	H	1		
Oxygen	O	2		

4. Select a ball representing a carbon atom. Insert a stick to indicate, "Have electron, will share." Put a hydrogen atom on the other end of the stick.
   (a) What kind of chemical bond holds the carbon atom and the hydrogen atom together? ______________________
   (b) Can the hydrogen atom form any more bonds? ______________ Explain. ________
   ____
   (c) Can the carbon atom form more than one chemical bond? ______________ Why do you think so? ______________________
5. Insert sticks into any remaining holes in the ball representing the carbon atom. Place hydrogen atoms on the other end of each stick. The model you have just made represents a molecule of methane gas.
6. *A STRUCTURAL FORMULA shows symbols for atoms connected by dashes to represent chemical bonds.* A dash between two atoms indicates that the atoms are bonded together by a covalent bond.

Write a structural formula to represent the methane molecule. Write the atom's symbol for each ball and a dash for each stick you have used.

7. Select from the kit, balls representing two carbon atoms, one oxygen atom, and six hydrogen atoms.
8. Attach the two carbon atoms to each other with a covalent bond. Add on all of the rest of the selected atoms to complete the model. You have made a model of ethanol.

Show what the model looks like by writing its structural formula.

MODEL OF ETHANOL:

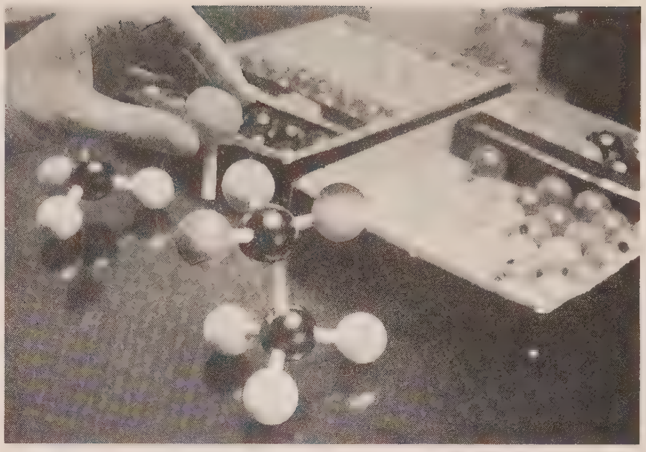
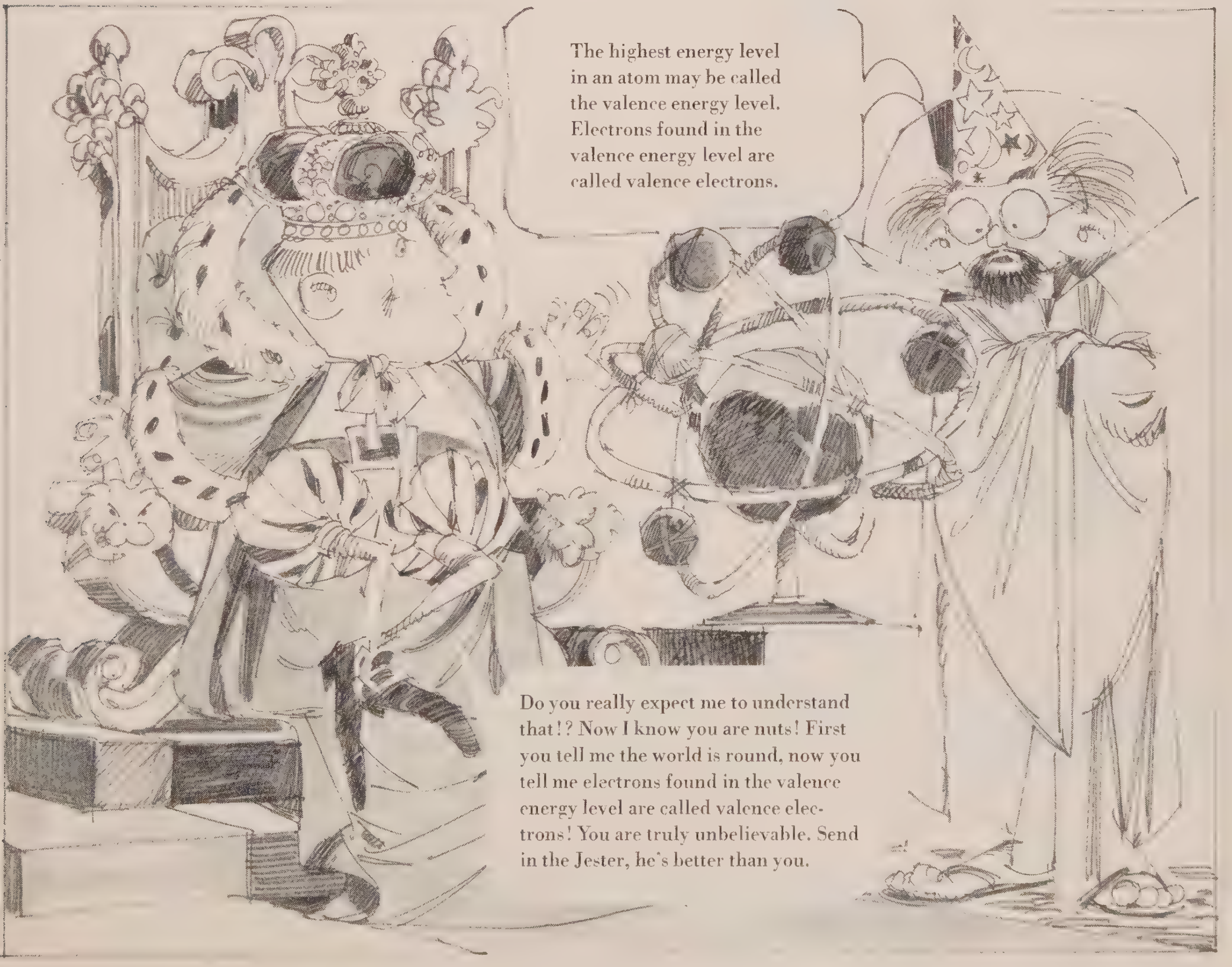


Fig. 47. *(Clara Aich and George Senty)*

9. Make a model of a water molecule, H_2O. Write its structural formula.

MODEL OF WATER:


The highest energy level in an atom may be called the valence energy level. Electrons found in the valence energy level are called valence electrons.
Do you really expect me to understand that!? Now I know you are nuts! First you tell me the world is round, now you tell me electrons found in the valence energy level are called valence electrons! You are truly unbelievable. Send in the Jester, he's better than you.

Questions to answer

1. How would you define a molecule? ______

   ______

   Do molecules always have the same number of atoms? ______

   ______

   Explain. ______

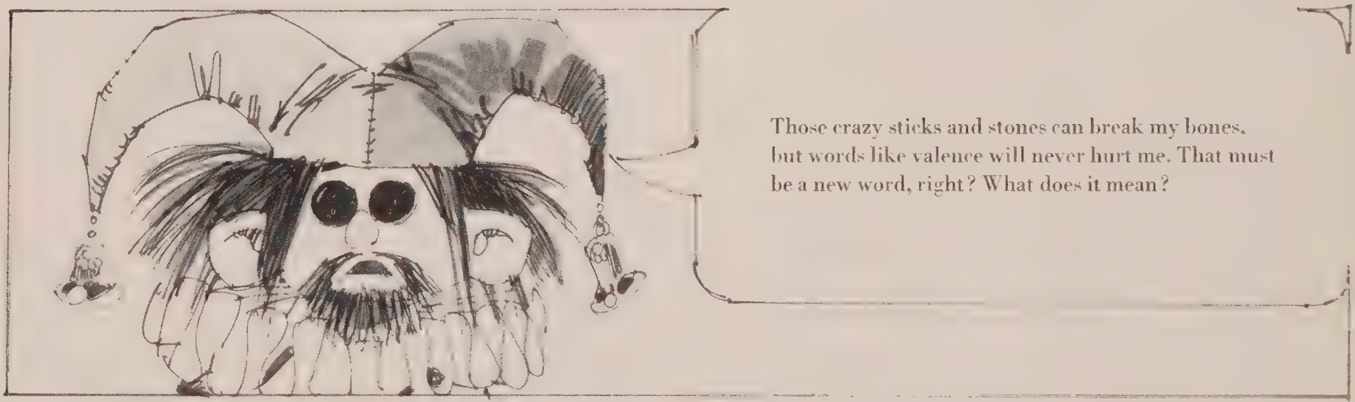
   ______

2. Oxygen can make two covalent bonds, but hydrogen can make only one. Explain. ______

   ______

3. Select from the list below the compounds which are most likely to have covalent bonds. (Refer to the introduction of this experiment.)

	Yes	*No*	*Why?*
Ammonia, NH_3			
Hydrogen chloride, HCl			
Sodium chloride, NaCl			
Silicon dioxide, SiO_2			
Potassium bromide, KBr			


Those crazy sticks and stones can break my bones, but words like valence will never hurt me. That must be a new word, right? What does it mean?

Extra for experts

What happens when the nucleus of one atom attracts a pair of shared electrons more strongly than the nucleus of the other atom? ______

______

______

______

______

______

Atom Hopping

You have learned that as you go across the Periodic Table from left to right that each succeeding atom has one more proton. To balance the additional proton there is one more electron. When electrons are added to the valence energy level as in our model of the atom, the electrons are added in the order shown in Fig. 48.

The completely-filled valence energy level has four pairs of electrons. In order to have the eight electrons, the atom forms a bond with another atom.

Valence electrons are the parts of the atom which form bonds. All other electrons in the atom, together with the nucleus, do not change. They make up what is known as the CORE of the atom.

You have studied covalent bonding in Experiment 10.1. In this experiment you will learn about electrically neutral atoms becoming charged particles called IONS. When the ions are formed, an IONIC BOND is also formed.

When one or more valence electrons are transferred from one atom to another, an IONIC BOND is formed.

What can happen if a valence electron is pulled away from an electrically neutral atom?

What can happen if a valence electron is added to an electrically neutral atom?

Pass me the three ion. I'm going to give this thing a charge!

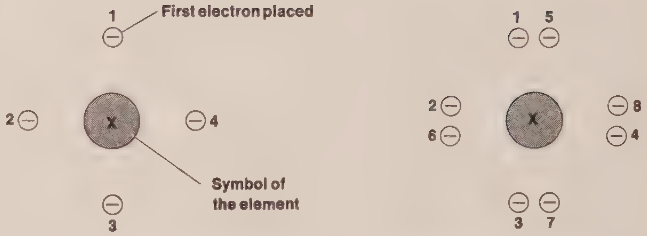
Experiment 10.2

Equipment and Supplies per Lab Station

sheet each red, green, and gray paper
(16) markers, possibly marble chips
compasses; scissors

What to do

1. Obtain small sheets of red, green, and gray colored papers. (Any color will do, just so you can tell the difference among them.) Out of each paper cut a circle with a radius of about two inches. Use a circle to represent the core of each atom for which you make a model.
2. Obtain marble chips or other small objects. Place the objects around the circles to represent valence electrons.
3. *Model of the sodium atom*
   Label a circle Na for sodium. Place one small marker next to the circle.
4. *Model of the chlorine atom*
   Label a circle Cl for chlorine. Place seven small markers around the circle. Follow the order shown in Fig. 48.
5. *Model for the neon atom*
   Label a third circle Ne for neon. Complete the model by placing eight small markers around the circle. Again follow the order in Fig. 48.
6. Study the Periodic Table to help you fill in the Data Chart for the sodium atom, the chlorine atom and neon atom.



A. Placing the first four electrons in the valence energy level.

B. Pairing of the electrons in the valence energy level.

Fig. 48.

DATA CHART

	Na atom	Na⁺ ion	Cl atom	Cl⁻ ion	Ne atom
Number of protons					
Number of electrons in core					
Total electric charge on the core					
Number of valence electrons					
Total electric charge					

7. Transfer the valence electron from the sodium, Na, model to the chlorine, Cl, model.
   (a) Does the Na model still represent an atom of sodium? ______________ Explain your answer. ______________________________
   ____
   (b) What happened to the neutral charge on the chlorine atom? __________________________
   ____

Is this model still a model for the chlorine atom? ____________ If not, what is it called? ____________________

(c) Complete the Data Chart.

8. Compare the three models.
   (a) How do the charges compare?

   Sodium ion ____________ Chloride ion ____________ Neon atom ____________

   (b) How many incomplete energy levels are present?

   in the sodium atom ____________

   in the sodium ion ____________

   in the neon atom ____________

   in the chlorine atom ____________

   in the chlorine ion ____________

   (c) Can you think of a reason why neon is considered to be relatively unattracted to other atoms? ____________________

   ____

9. When one atom transfers an electron to another atom, the two atoms may form a chemical bond. What is the name for this type chemical bond?

   ____

   ____

Questions to answer

1. (a) How many valence electrons would a magnesium, Mg, atom have to lose to become an ion with an outermost energy level containing eight electrons? ____________

   Explain. ____________________

   ____

   (b) How many chlorine, Cl, atoms would you use to take the valence electrons from a magnesium atom? ____________

2. How many valence electrons would an oxygen atom have to gain to become an oxygen ion with an outermost energy level containing eight electrons? ____________ Explain. ____________

   ____

   ____

Yeah, why? Show me! Go ahead show me!

3. State a rule to be used in changing atoms
   (a) into positive ions. ____________________

   ____

   (b) into negative ions. ____________________

   ____

4. (a) How can a positive ion become a neutral atom again? ____________________

   ____

   (b) How can a negative ion become a neutral atom again? ____________________

   ____

NUMBER, PLEASE

In this experiment you will find the least number of atoms which can combine by transferring electrons from atoms of metals to atoms of nonmetals. The chemical bond formed in this way is an ionic bond.

There are some rules to think about.

1. All of the *valence* electrons of each metal atom are transferred out of the atom. The positive charge of the ion formed is equal to the number of electrons transferred out.
2. Electrons are added to nonmetal atoms until each atom has eight *valence* electrons. The negative charge on the ion formed is equal to the number of electrons added.
3. When metal atoms and nonmetal atoms unite in this way, the *number of electrons transferred out of the metal atom must equal the number of electrons added to the nonmetal atoms.*
4. The number of electrons transferred from a metal atom is known as the *combining number* of the metal atom. It is written as plus, +, representing positive electric charges. The number of electrons added to a nonmetal atom is known as the *combining number* of the nonmetal atom. It is written as minus, –, representing negative electric charges. *The number and kinds of atoms combining chemically depends upon the combining numbers of the atoms.*
5. The *chemical formula* is a shorthand way of expressing the kinds and numbers of atoms chemically united.
    - (a) The symbols of the atoms in a formula show what atoms are united. Example: $NaCl$ means that sodium and chlorine atoms are united.
    - (b) *Subscript numbers* are numbers written at the lower right corner of the symbol for the atom in a formula. The subscript number tells the number of atoms present in the combination. If there is no subscript number, one is understood. Example: Li_2O means that two atoms of lithium and one atom of oxygen are united chemically.

(c) *Superscript numbers* are numbers of plus, +, or minus, −, signs written at the top right of the *symbol for an ion.* The number of plus, +, signs tells the number of electrons transferred out of the atom. Examples: Na^+, Ca^{++}, Al^{+3}. The number of minus, −, signs shows the number of electrons added to the atom. Example: Br^-, O^{-2}. *The formula for an ion* is the symbol for the atom with the correct number of plus or minus signs.

(d) The combining number of an atom, plus or minus, is written in a formula only as an aid in finding the correct number of atoms to make the positive and negative charges equal. Calculation of the number of atoms:

combining number of the +ion × the subscript = total + charge
combining number of the −ion × the subscript = total − charge

---

The TOTAL + plus THE TOTAL MUST = ZERO
Example: What is the correct formula for a compound formed from K^+ and O^{-2}?

$1+ \times 2 = 2+$
$\underline{2- \times 1 = 2-}$

$2+ \text{ plus } 2- = 0$
The formula is K_2O $(K_2^{+}O_1^{-2})$

Experiment 10.3

Equipment and Supplies per Lab Station
compass; scissors
(20) markers for electrons
two differently-colored papers

What to do

1. Obtain several circles of colored paper. Write the letter *M* on them. These circles will represent the core of metal atoms.
2. Obtain several circles of another color. Write the letter *N* on them. These circles will represent the core of the nonmetal atoms.
3. Make models of the metallic magnesium atom. You may use as many of the models as you need when you do step 5.
4. Make models of the nonmetallic chlorine atom. You may use as many of these as are necessary in doing step 5.
5. Transfer the valence electrons from a model of the magnesium atom to one or more models of the chlorine atom. NOTE: EVERY MAGNESIUM ATOM YOU USE MUST LOSE BOTH VALENCE ELECTRONS. EVERY CHLORINE ATOM MUST END WITH EIGHT VALENCE ELECTRONS. Why? ________________

____

In your model, how many magnesium atoms were used? ________________

How many chlorine atoms were used? ________

____

6. Complete the formula for the compound in step 5 by writing the least number of atoms used at the lower right corner of each symbol.
Mg ______ Cl ______

7. (a) Indicate the combining number of magnesium. Mg———

   (b) Indicate the combining number of chlorine.
   Cl———

8. Repeat steps 1 through 7, using sodium in place of magnesium and sulfur for the nonmetal. Complete the formula below by adding the signs to indicate the combining number of each atom and the subscript numbers to show the least number of atoms used. $Na^{\_\_\_\_}_{\_\_\_\_}$ $S^{\_\_\_\_}_{\_\_\_\_}$

9. Write the symbol for an atom of aluminum. Place its combining number at the top right corner of the symbol. Write the symbol for an atom of chlorine next to the aluminum. Place chlorine's combining number at the top right of the symbol. Write the least number of atoms you think each symbol should have.

$Al^{\_\_\_\_}_{\_\_\_\_}$ $Cl^{\_\_\_\_}_{\_\_\_\_}$

Check your formula: The sum of the total of the total combining numbers must equal zero.

combining number of Al × atoms of Al = total +
combining number of Cl × atoms of Cl = total −
TOTAL + plus TOTAL − MUST = ZERO

If your formula does not give this sum, then change the subscripts until the zero is reached.

Questions to answer

1. *Check* the incorrect formulas in the following list. Write the correct formula.

	Incorrect	Correct		Incorrect	Correct
MgO			NaS		
H_2O			AlS_3		
HCl_2			$CaCl_2$		
$AlCl_2$					

2. Write the formula for each of the following:

Zinc sulfide ______________ Zinc oxide ______________

Zinc chloride ______________

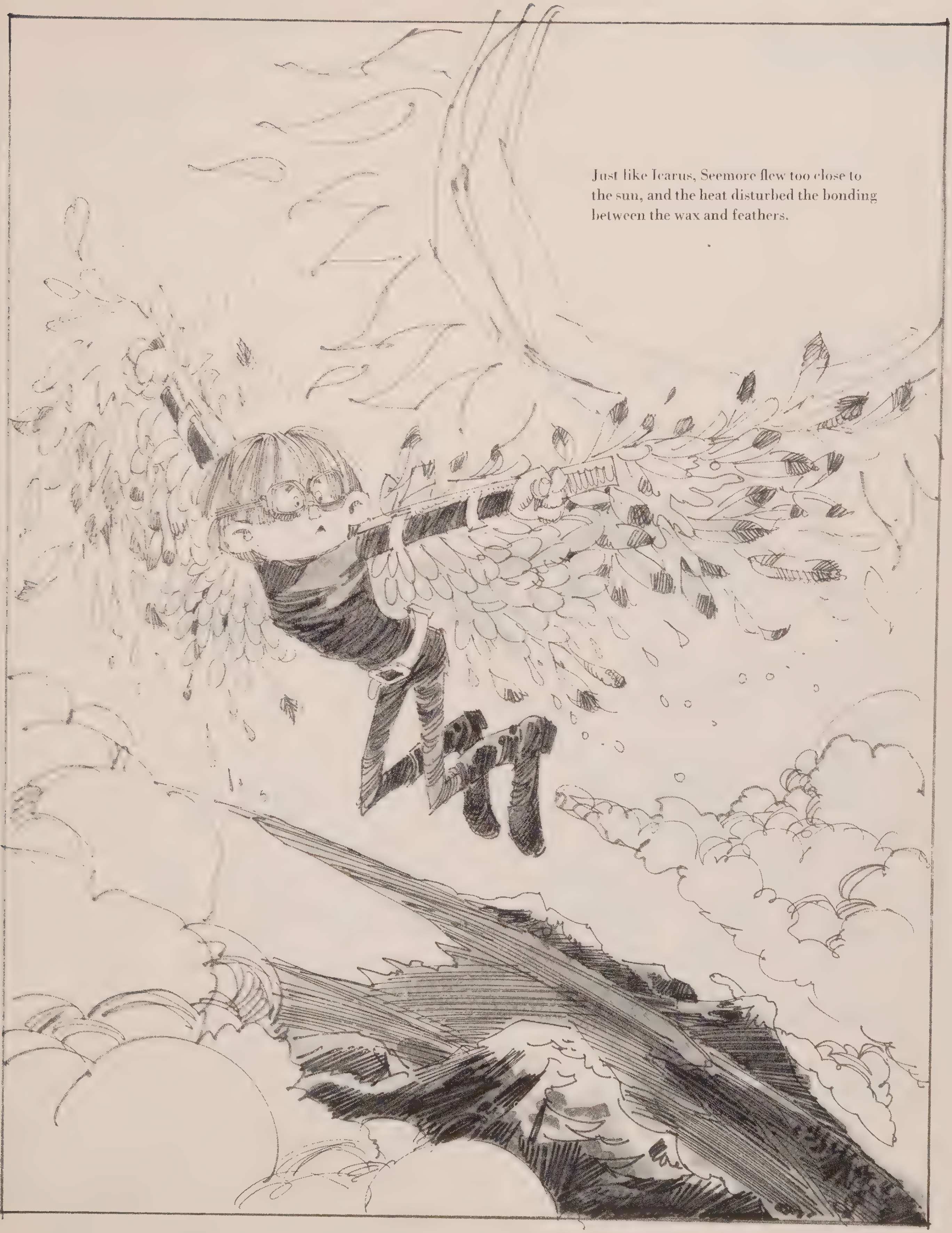
3. Name each of the following:

$BaBr_2$ ______________ MgI_2 ______________

Extra for experts

Using the rules that metals lose all valence electrons, and nonmetals build up to eight valence electrons, see if you can find the least number of atoms of aluminum and oxygen which are needed to write the formula for aluminum oxide. Use models if you wish. The formula is ______________

____



Just like Icarus, Seemore flew too close to the sun, and the heat disturbed the bonding between the wax and feathers.

Exercise

Writing Electron – Dot Formulas

The models of atoms and ions which you used in Experiment 10.2 may be shown by electron-dot formulas.

PART I FORMULAS FOR ATOMS

1. Write the electron-dot formula for an atom of sodium by:
   (a) writing the symbol for sodium to represent the core of the atom.
   (b) placing a dot representing a valence electron beside the symbol.
   The electron-dot formula for a sodium atom is ____________________________

2. (a) Refer to your Periodic Table. Atoms of Period 2 have numbers of electrons in the valence energy level as follows:

Li)	Be)	B)	C)	N)	O)	F)	Ne)
1e$^-$	2e$^-$	3e$^-$	4e$^-$	5e$^-$	6e$^-$	7e$^-$	8e$^-$

   (b) Refer to Fig. 48 in Experiment 10.2 for the placing of valence electrons about the core of the atom as they are thought to be arranged at the time of chemical reaction.
      (1) The electron-dot formulas of the first four of the Period 2 atoms may be represented as follows:

$\mathrm{Li}\cdot$ $\quad$ $\overset{\cdot}{\mathrm{B}}\mathrm{e}\cdot$ $\quad$ $\underset{\cdot}{\overset{\cdot}{\mathrm{B}}}\cdot$ $\quad$ $\cdot\underset{\cdot}{\overset{\cdot}{\mathrm{C}}}\cdot$

      (2) Pairing of electrons starts with nitrogen as shown below. Complete the electron-dot formulas of the last three atoms.

$\cdot\underset{\cdot}{\overset{\cdot\cdot}{\mathrm{N}}}\cdot$ $\quad$ O $\quad$ F $\quad$ Ne

   (c) Write the electron-dot formula for atom X which has four valence electrons. __________
   (d) Write the electron-dot formula for an atom Y which has seven valence electrons. ______

3. Write the electron-dot formulas for an atom of each of the elements listed in Period 3 of the Periodic Table.

Na Mg Al Si P S Cl Ar

4. Write the electron-dot formulas for atoms in the following families of elements:

Group I A	Group II A	Group VII A
Li	Be	F
Na	Mg	Cl
K	Ca	Br
Rb	Sr	I

PART II FORMULAS FOR IONS

5. Study the electron-dot formulas for some positive ions listed below:
Electron-dot formulas for the atoms $\mathrm{Li}\cdot$ $\dot{\mathrm{Be}}\cdot$ $\underset{\cdot}{\dot{\mathrm{B}}}\cdot$

Formulas for the ions Li^{+} Be^{++} B^{+3}

(a) How do the electron-dot formulas for the atoms and the ions of these elements differ?

____

____

(b) How was the positive charge produced? ______

____

(c) State a rule for writing electron-dot formulas for ions of atoms having 1, 2, or 3 valence electrons. ______________

____

6. Study the electron-dot formulas for ions of atoms having 5, 6, or 7 valence electrons respectively.

Electron-dot formulas for the atoms $\cdot\overset{\cdot\cdot}{\underset{\cdot}{\mathrm{N}}}\cdot$ $\cdot\overset{\cdot\cdot}{\underset{\cdot}{\mathrm{O}}}:$ $\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Cl}}}:$

Formulas for the ions $\left[:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{N}}}:\right]^{-3}$ $\left[:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{O}}}:\right]^{-2}$ $\left[:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Cl}}}:\right]^{-1}$

(a) How are the formulas alike? ______________

____

(b) How are the formulas different? ______________

____

(c) How did the ions get the electric charge?

____

____

7. Neon has eight valence electrons. Do you think neon can readily become an ion? ______________

Why or why not? ______________________

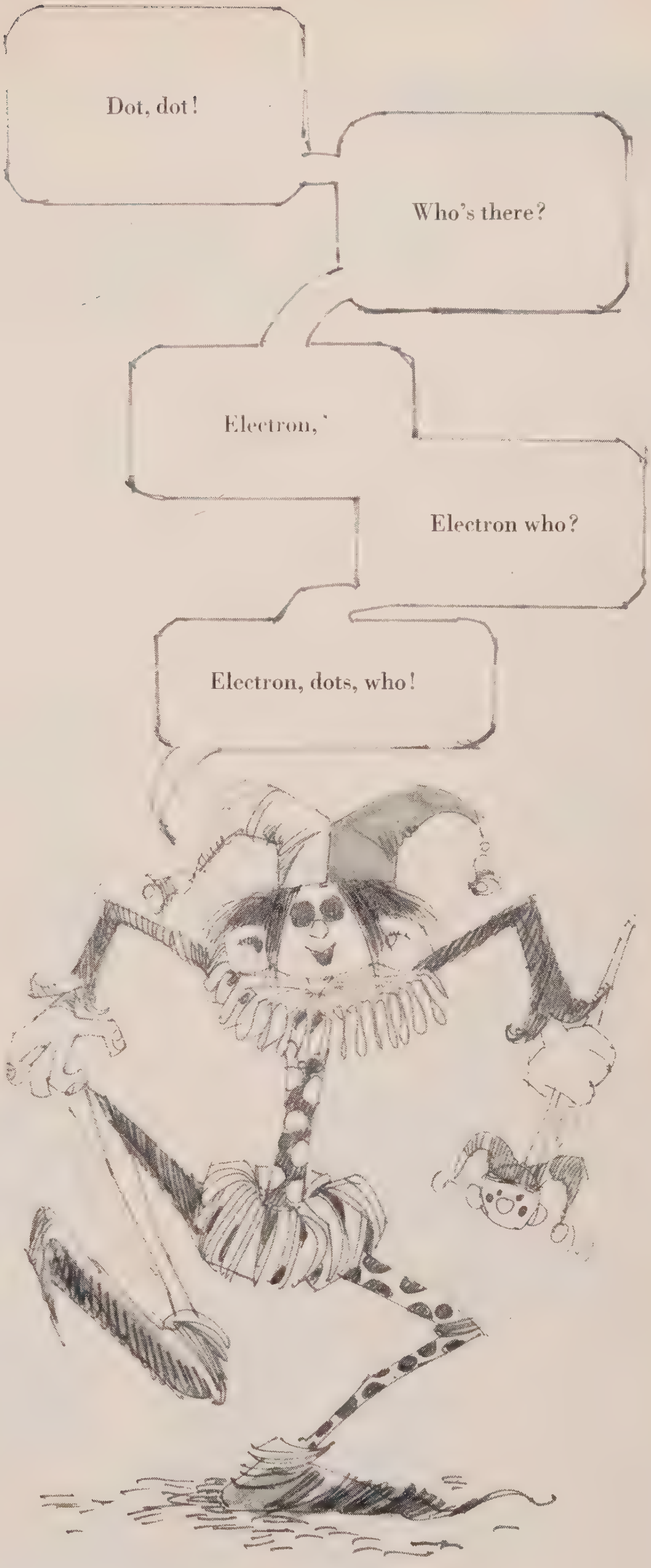
____

PART III ELECTRON-DOT FORMULAS FOR COMPOUNDS

8. Write the electron-dot formula for sodium chloride by:

(a) writing the electron-dot formula for a sodium ion.

(b) writing the electron-dot formula for a chloride ion beside the sodium ion.


Dot, dot!
Who's there?
Electron,
Electron who?
Electron, dots, who!

The formula for sodium chloride is ______________

9. In a similar manner, write the electron-dot formulas for

(a) Calcium chloride ______________

(b) Magnesium sulfide ______________

(c) Aluminum chloride ______________

(d) Aluminum sulfide ______________

Exercise

WRITING FORMULAS

Electron-dot formulas are written when you are working out ions and when you are trying to find the correct formula for a compound. The formula for a compound, however, is made up only of the symbols of the atoms combined and the subscript numbers.

Write formulas for the following compounds:

1. Potassium chloride ____________________
2. Magnesium bromide ____________________
3. Sodium fluoride ____________________
4. Barium chloride ____________________
5. Aluminum bromide ____________________
6. Calcium oxide ____________________
7. Strontium iodide ____________________
8. Lithium sulfide ____________________
9. Aluminum sulfide ____________________
10. Rubidium chloride ____________________

Name the following compounds:

1. Al_2O_3 ______________________________
2. $BaBr_2$ ______________________________
3. KI ______________________________
4. $CaCl_2$ ______________________________
5. Na_2S ______________________________

WHAT'S ITS NATURE?

Some properties may be used to predict the type of bonding in a substance. These properties are: phase at room temperature, melting point, solubility in water, and electrical conductivity.

In this experiment you will find how these properties vary in ionic and covalent bonded substances.

I wonder if he knew about covalent and ionic bonds?

Equipment and Supplies per Lab Station

(9) beakers to fit conductivity apparatus
(9) test tubes; (9) stoppers to fit test tubes
test tube rack; balance; conductivity apparatus
distilled water; oil; ethanol; starch; glycerine; sucrose
calcium chloride (anhydrous); sodium chloride
potassium iodide

What to do

1. Inspect each substance listed in the Data Chart. Record the phase of each at room temperature.
2. Note the melting point of each substance listed in the Data Chart.
3. Your teacher will demonstrate and explain the conductivity apparatus. See Fig. 49.
   (a) Under what conditions will the bulb in the apparatus light? ______________________

   (b) If the bulb lights, what property must a substance between the electrodes have? ___

   ____
4. Record in the Data Chart what happens as the electrodes are touched in turn to each substance.
5. Cover each of the substances tested in step 4 with distilled water and stir. Then touch the electrodes to each substance again. Record your observations in the Data Chart.
6. Put approximately 1 gram of one of the substances into a test tube. Add 10–15 ml of water to the tube and shake. Note the apparent solubility of the substance.

DATA CHART

Substance	Type of Bonding	Phase at Room Temp.	Melting Point °C	Electrical Conductivity		Solubility
				Without Water	With Water	
Distilled water	covalent		0			
Sodium chloride	ionic		801			
Potassium iodide			686			
Sucrose			185			
Oil			less than 20*			
Ethanol			−117.3			
Starch	covalent		decomposes			
Glycerine			18.6			
Calcium chloride			772			

*varies with type of oil

7. Repeat step 6 with each of the other substances and write your observations in the Data Chart.
8. Sodium chloride has ionic bonding. Study the chart to find one or more substance(s) which has (have) similar properties.
   (a) What substance(s) would you classify as having ionic bonds? ______________________

   ______________________________ .

   (b) Why did you select these substances?

   ____

   ____
9. Oil has covalent bonds. What other substances tested would you say have covalent bonding?

   ____

   Why did you select these substances? __________

   ____

   ____
10. Water is a polar substance and can dissolve other polar substances. Which substances of those tested do you think are polar? __________

    ____

    Explain. ______________________________

    ____

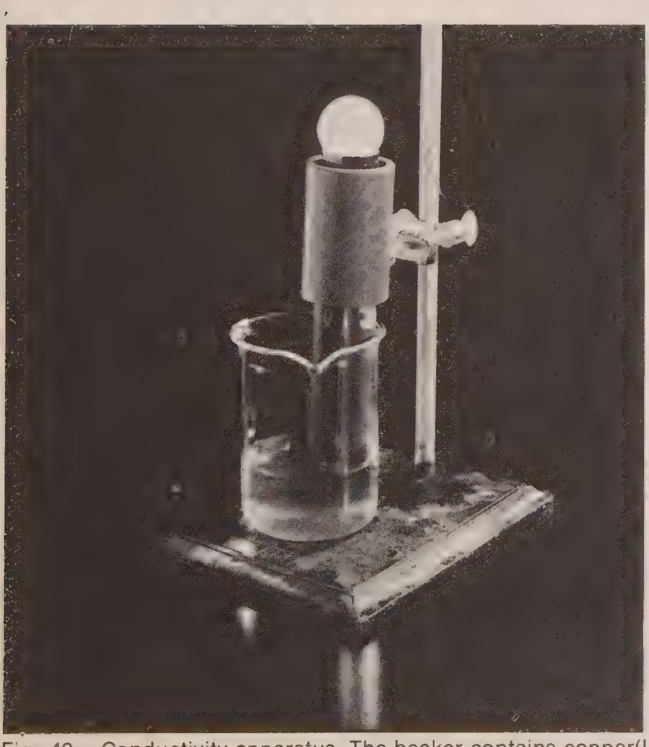


Fig. 49. Conductivity apparatus. The beaker contains copper(II) sulfate. *(Clara Aich and George Senty)*

Questions to answer

1. What general properties do covalent-bonded substances have? ______________________

   ____

   ____
2. What general properties do ionic-bonded substances have? ______________________

   ____

   ____

Chapter 11

Exercise

SCIENTIFIC NOTATION

Expressing a count of atoms large enough to weigh on a laboratory balance calls for a very large number. Expressing the mass of a single atom in grams requires a very small number. To simplify the writing of these large and small numbers, chemists use "scientific notation." The following examples will show you how this is done.

PART I VERY LARGE NUMBERS

1. Write 12.0 in scientific notation.
   (a) Move the decimal point to the *left* until only one numeral stands in front of the decimal point.

$$\underset{\curvearrowleft}{1\,2.0} = 1.2$$

   You have *divided* 12.0 by 10.
   (b) You moved the decimal point *one* place to the *left*. Write the one as a small number to the right and slightly above 10, that is, 10^1.
   (c) Write 12.0 in scientific notation as *1.2×10^1*. This means that 1.2 *multiplied by 10 once equals 12.0.*
2. Write 2542.0 in scientific notation.
   (a) Move the decimal point to the *left* until only one numeral stands in front of the decimal point.

$$\underset{\longleftarrow}{2\,5\,4\,2.0} = 2.542$$

   You have *divided* 2542.0 by 1000.

"There's got to be an easier way to write how many atoms are on the head of a pin."

   (b) Count the number of places you moved the decimal point. In this problem you moved it three places. Dividing by 1000 is the same as dividing by 10 three times.
   (c) Write the three to the right and slightly above the 10, that is, 10^3.
   (d) Write 2542.0 in scientific notation as 2.542×10^3. This means that *2.542 is multiplied by 10 three times.*
   (e) Numbers expressed in scientific notation may be rounded off so that there are only two numerals after the decimal point. In 2.542, the second 2 is less than 5. It may be dropped. If so, write the number as 2.54×10^3. Suppose a numeral between 5 and nine was after the 4 in the above—for example 2.545. Then you round the second number, after the decimal point, to the next higher number. The number is now written 2.55×10^3.

PART II VERY SMALL NUMBERS

1. Write 0.625 in scientific notation.
   (a) Move the decimal point to the *right* until there is only one numeral in front of the decimal point.

$$0.6\,2\,5 = 6.25$$

You have *multiplied* 0.625 by 10.

   (b) You have moved the decimal point one place to the *right*. Write the one as a *negative* number to the right and slightly above the 10, that is, 10^{-1}.

   (c) Write 0.625 in scientific notation as 6.25×10^{-1}. This means that *6.25 divided by 10 once equals 0.625.*

2. Write 0.00072 in scientific notation.
   (a) Move the decimal point to the *right* until there is only one numeral in front of the decimal point.

$$0.0\,0\,0\,7\,2 = 7.2$$

You have *multiplied* 0.00072 by 10,000.

   (b) Count the number of places you moved the decimal point. In this problem you moved it four places. Multiplying by 10,000 is the same as multiplying by 10 four times.
   (c) Write the four as a *negative* number to the right and slightly above 10, that is, 10^{-4}.
   (d) Write 0.00072 in scientific notation as 7.2×10^{-4}. This means that 7.2 is *divided* by 10 four times.

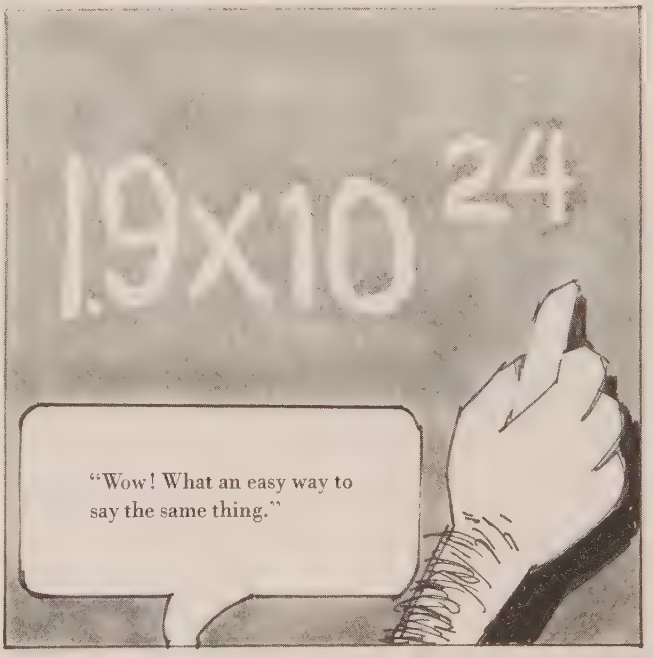
Problems

Express the following numbers in scientific notation.

1. 225 ____________
2. 4672 ____________
3. 8,200,000,000 ____________
4. 961 ____________
5. 5,620 ____________
6. 0.23 ____________
7. 0.542 ____________
8. 0.0082 ____________
9. 0.000045 ____________
10. 0.0000007 ____________

Express the following numbers as ordinary numbers.

1. 5×10^{2} ____________
2. 2.6×10^{-2} ____________
3. 8.64×10^{-4} ____________
4. 9.52×10^{-14} ____________
5. 7.2×10^{18} ____________


1.9X10 24
"Wow! What an easy way to say the same thing."

Extra for experts

Scientific notation is also know as the "powers of 10."

A *positive* number to the right and slightly above the 10 is a *positive power of 10*. It shows how many times the number in the expression is multiplied by 10.

A *negative* number to the right and slightly above 10 is a *negative power of 10*. It shows how many times the number in the expression is divided by 10.

1. The Avogadro Number is 6.02×10^{23}.
   (a) What is the *power of 10* in this expression?

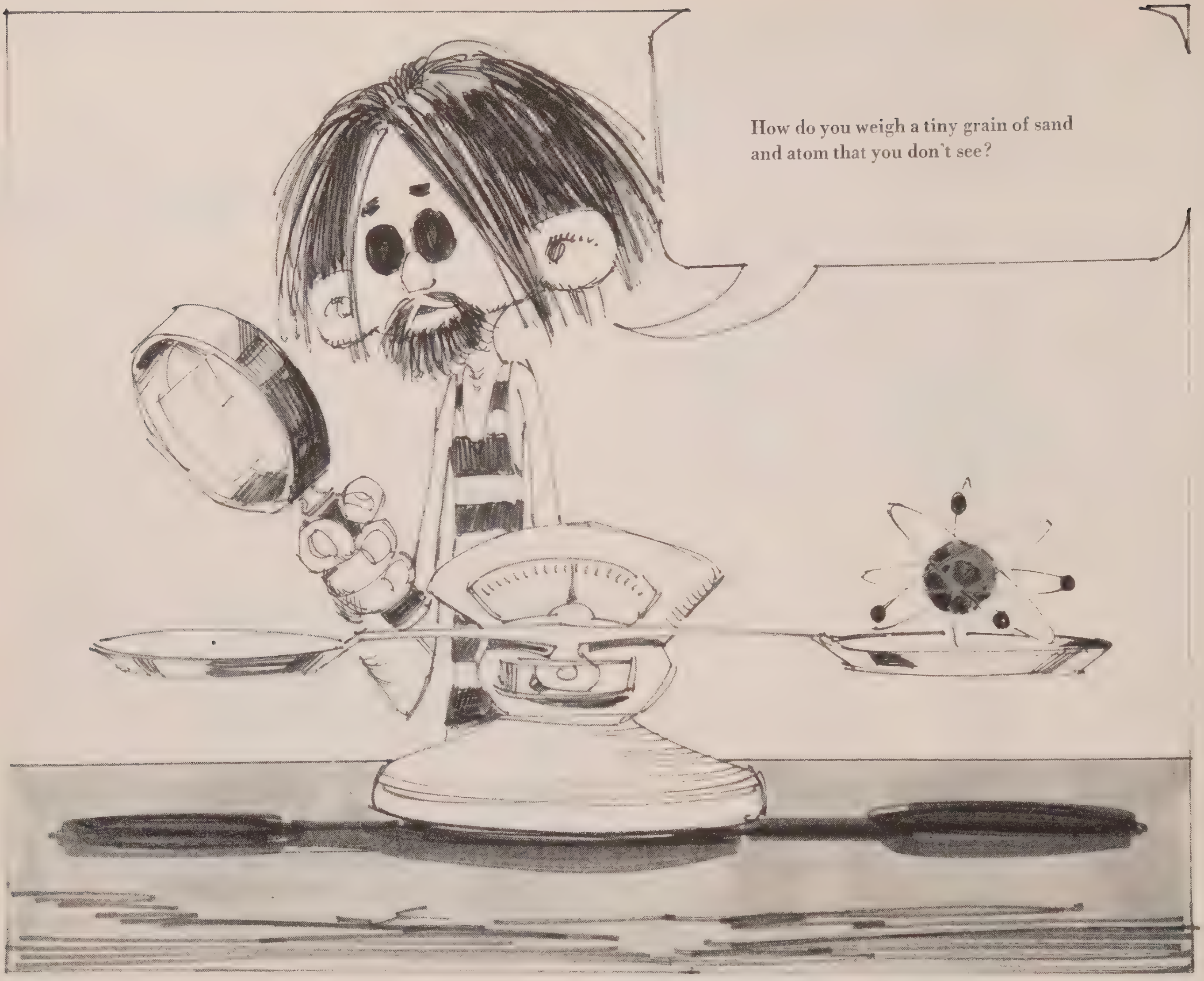
   ____

   (b) How many zeroes would you have to write after 602 to write 6.02×10^{23} as an ordinary number? ____________

2. The mass of a proton is 1.672×10^{-24}.
   (a) What is the *power of 10* in this expression?

   ____

   (b) How many zeroes would you have to write to the left between 1672 and a decimal point to write 1.672×10^{-24} as an ordinary number? ____________


How do you weigh a tiny grain of sand
and atom that you don't see?

A GRAIN OF SAND

Grains of sand are very small. They can be seen, but are difficult to weigh. In this experiment you will learn how to determine the mass of a grain of sand. Atoms are much smaller than grains of sand, but you will find that their masses can also be determined.

Experiment 11.1

Equipment and Supplies per Lab Station

watch glass; beaker—150 ml; glass stirring rod
sharp pin, size used for tack board
forceps, if grains of sand are large enough to grasp
teaspoonful of sand, 30 mesh or coarser
sheet of graph paper, notebook size, squares 1 cm × 1 cm

What to do

1. Obtain about one teaspoonful of coarse sand on your watch glass.
2. Weigh a small beaker. Leave the beaker and its weight in position on the balance.
3. Carefully transfer a single grain of sand to the beaker. The grain of sand may be pushed from the watch glass into the beaker with the sharp point of a pin. Can you see any change in the balance reading? ____________________
4. Add nine more grains of sand to the beaker. Do the ten grains make any difference in the balance reading? ____________________
5. Adjust the balance reading to be one gram more than the mass of the beaker. Add sand until you have *one gram* of sand in the beaker.
6. Obtain a piece of graph paper. The squares should be one centimeter by one centimeter.
7. Spread the one gram of sand evenly over the graph paper. Be careful not to lose any. Rolling a glass stirring rod over the sand will help to spread it evenly.
   How many squares are covered with sand?
   ____
8. Pick out one of the squares. Brush all of the sand off all of the squares except this one. Try not to disturb the grains on the chosen square.
9. Count the total number of grains of sand on your chosen square of graph paper. Use the point of the pin to move each grain out of the square. See Fig. 50.
   Number of grains of sand on chosen square =
   ____
10. Calculate the number of grains of sand in the one gram of sand: grains of sand in one square *TIMES* number of squares covered *EQUALS* grains of sand in one gram of sand.
11. Change the very large number of grains of sand in one gram to scientific notation as you changed the numbers in Exercise: *Scientific Notation.*

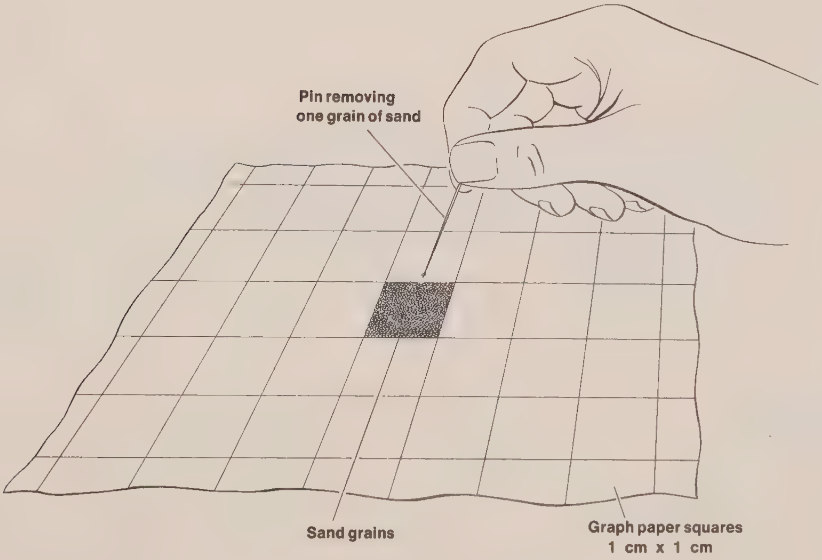

Pin removing one grain of sand
Sand grains
Graph paper squares 1 cm x 1 cm

Fig. 50.

________________ grains of sand = ________________ grains of sand
(in one gram) (scientific notation)

12. Calculate the mass of a single grain of sand. Divide one gram of sand by the number of grains in one gram of sand.

$\frac{\text{1 gram of sand}}{\text{grains of sand in one gram}} =$ ________________g per one grain of sand

13. Change this very small number of grams per one grain to scientific notation as you changed the numbers in Exercise: *Scientific Notation.*

________________ grams per grain = ________________ grams per grain
(of sand) (scientific notation)

The grams per grain of sand is the mass of one grain of the sand.

14. Summarize your results as follows:
    (a) Number of squares on the graph paper covered by the one gram of sand ________
    ____
    (b) Number of grains of sand on one square
    ____
    (c) Grains of sand in one gram of sand ________
    ________________ (Express in scientific notation)
    (d) Mass in grams of one grain of sand ________
    ________________ (Express in scientific notation)

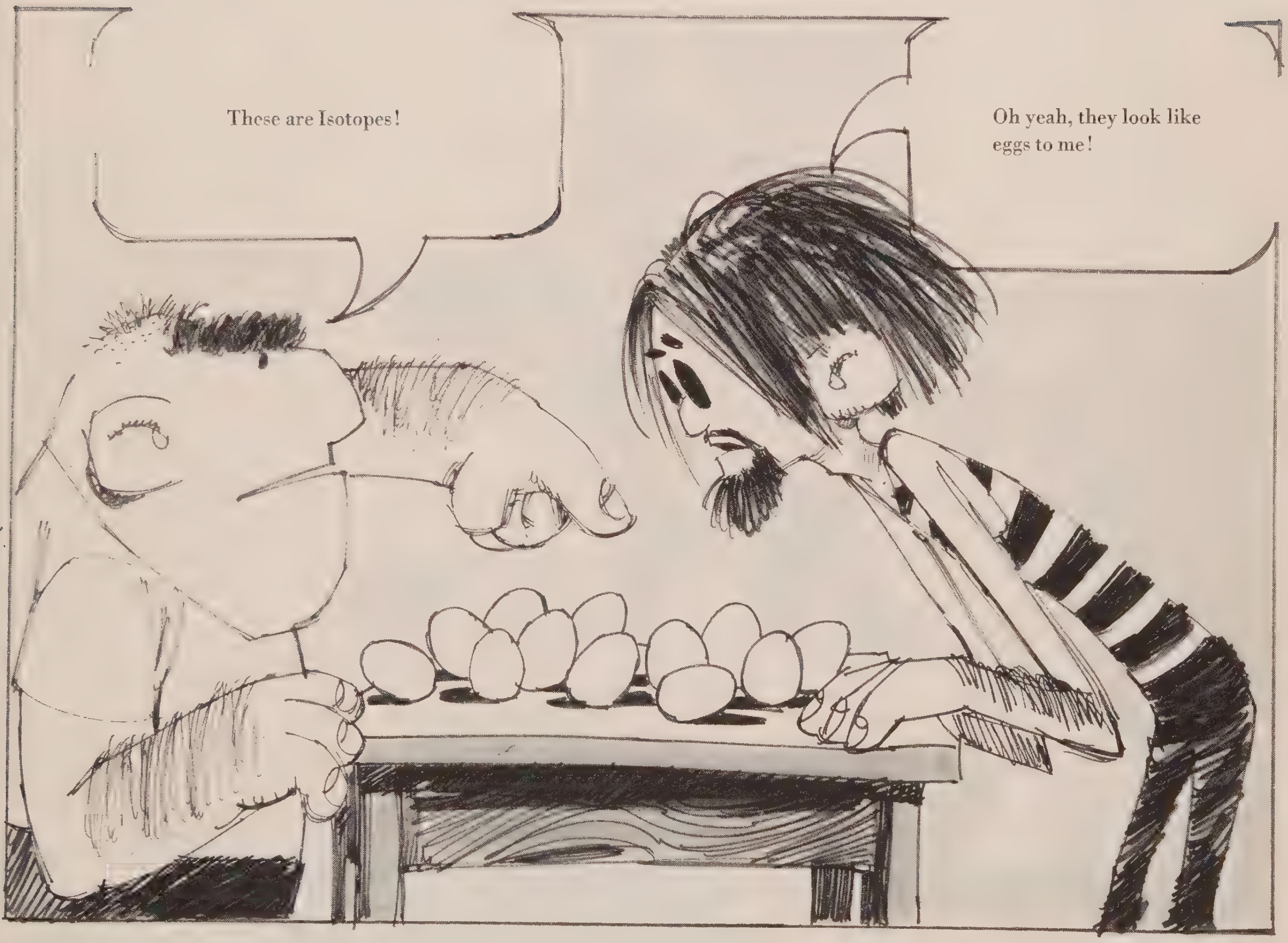
Questions to answer

1. By what three steps did you find the mass in grams of one grain of sand?
    (a) ________________________________
    (b) ________________________________
    (c) ________________________________
    ____
2. Since atoms are so small, larger numbers of them than of grains of sand are needed if they are to be weighed on a laboratory balance.
    You cannot count atoms as you counted the grains of sand, but scientists have counted atoms in solids with the help of X-rays.
    Suppose a chemist wishes to find the mass of an individual atom of an element, say copper.
    (a) He weighs exactly 63.5 grams of copper metal.
    (b) He finds that it has been determined with the help of X-rays that there are 6.02×10^{23} atoms of copper in 63.5 grams of solid copper.
    (*Note*: 6.02×10^{23} atoms of any element can be weighed on a laboratory balance.)
    (c) How will the chemist use this information to find the mass of one atom of copper? (You do not need to find the numerical answer.)
    ____
    ____
    ____

EGGS AND ATOMS

What is an isotope? What does it mean to say that the atoms in a sample of an element are isotopes of each other?

Ordinary eggs are a lot bigger than atoms, but perhaps they can give you one or two clues about isotopes.


These are Isotopes!
Oh yeah, they look like eggs to me!

Demonstration 11-A

Equipment and Supplies per Lab Station
beaker—250 ml; balance; 1 dozen eggs

What to do

1. Examine a dozen eggs in their box.

   Do the eggs appear to be alike? ____________

   If not, what differences do you see? ____________

   ____

2. Weigh each egg in turn in a weighed beaker and record its mass in grams.

   #1 ________g

   #2 ________g

   #3 ________g

   #4 ________g

   #5 ________g

   #6 ________g

   #7 ________g

   #8 ________g

   #9 ________g

   #10 ________g

   #11 ________g

   #12 ________g

   What is the total mass of all twelve eggs?

   ____________g

3. Calculate the average mass of the individual eggs. Divide the total mass of the twelve eggs by twelve.

   $$\frac{\text{total mass in grams of 12 eggs}}{12} =$$

   ____________g, ave. per egg

4. (a) Were any of the eggs definitely heavier than the average? ____________________

   ____

   (b) How many eggs were close to the average?

   ____

   ____

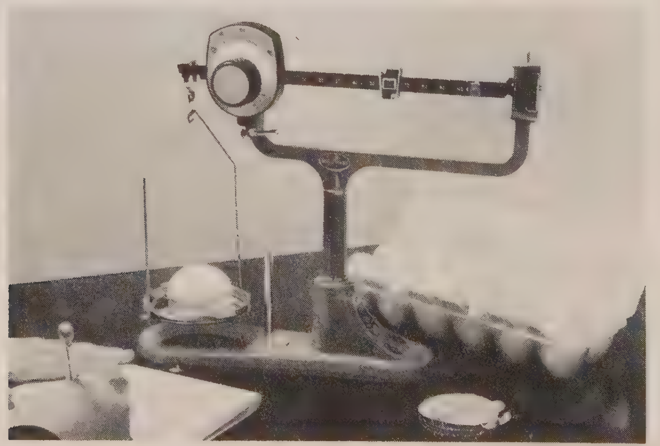


Fig. 51. Student is recording the mass of an egg. *(Clara Aich and George Senty)*

Questions to answer

1. The eggs you examined might be thought of as *isotopes* of each other. Let's see what this means.
   (a) How did you find the eggs to be all alike?

   ____

   ____

   (b) What measurement did you make that was different for some of the eggs? ____________

   ____

   (c) Then if the eggs are isotopes of each other, what are isotopes? ____________

   ____

   (d) Was the average mass of the eggs in the dozen a whole number or did it have a decimal fraction? ____________

   ____

2. The atoms in any naturally occurring sample of any element are said to be isotopes of each other. Let's see why.
   (a) What particle in the nucleus of the atom identifies the element? ____________

   (b) Neutrons are electrically neutral particles also present in the nuclei of atoms. Some atoms of the same element may have more neutrons in the nucleus than other atoms of the same element. A proton has one atomic mass unit. A neutron also has one atomic mass unit.

   Can the atomic mass of all atoms of an element be the same, if all the atoms of the elements do not have the same numbers of neutrons? ____________

   (c) How have you said above that all atoms of the same element are alike? ____________

   ____

   In what measurement may the atoms of an element (like the eggs) be different?

   ____

   ____

   Why may the atomic mass for the atoms of an element have a decimal fraction?

   ____

   ____

Completing The Periodic Table

You have found in your textbook that the atoms of each element have an average number of atomic mass units, amu. *One ATOMIC MASS UNIT is one twelfth the mass of the most abundant isotope of carbon which has an atomic mass of 12.011.*

The atomic mass units, amu, for an element, in terms of grams, are the gram atomic mass, grams per mole, for the element.

Exercise

What to do

Write the atomic mass units of each element in your Periodic Table, in the top right hand corner of the square for the element.

Questions to answer

1. In the Demonstration: *Eggs and Atoms*, what explanation did you find for the fact that the atomic mass units for most elements have decimal fractions? ______________________________
   ____
2. The number of neutrons in the nucleus of the atoms of average atomic mass in an element can be found by subtracting the atomic number from the rounded off atomic mass.

Element	Atomic Mass (rounded off)	Atomic number	Number of neutrons in the nucleus
Helium	4	2	2
Carbon	12	6	6
Sodium	23	11	12
Chlorine	35	17	18

Complete the following chart.

Element	Atomic mass (rounded off)	Atomic number	Number of neutrons in the nucleus
Fluorine	19		10
Sulfur	32	16	
Aluminum	27		14
Lithium		3	4
Argon		18	22

Finding The Number Of Moles In A Sample

The word "mole" has several meanings. In chemistry it stands for a count of objects just as a dozen or a gross stands for a count of objects.

1 dozen = 12 objects
1 gross = 144 objects
1 mole = 6.02×10^{23} objects or particles

One dozen objects can be weighed. At the market you might find that one dozen oranges weighed five pounds. You could write down "oranges — 5 lbs per dozen".

In chemistry the objects or particles in a mole are usually atoms, ions, or molecules. A mole of atoms or molecules can be weighed on an ordinary laboratory balance. *The MOLE MASS is the mass in grams of 6.02×10^{23} atoms or molecules.* It may be recorded as grams per mole (g/mole).

The mole mass may exist in three ways:

1. The gram atomic mass of an element
   Example: The gram atomic mass of sodium, Na, is 22.99 grams. This mass of sodium contains 6.02×10^{23} atoms.
2. The gram molecular mass of a molecular compound
   Example: The gram molecular mass of water, H_2O, is 18.00 grams. This mass of water contains 6.02×10^{23} H_2O molecules.

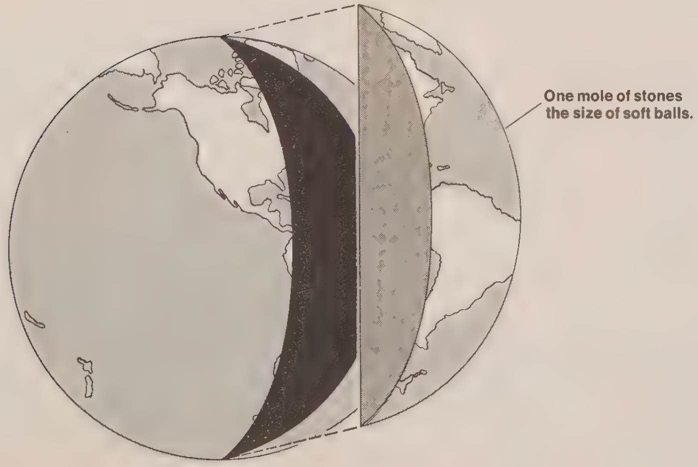

One mole of stones the size of soft balls.

Fig. 52. If the earth were made of stones the size of soft balls, it would take a mole of the stones to fill this space.

3. Gram formula mass of a compound
   Example: The formula of a compound which does not consist of molecules. It represents the ratio of the atoms in the compound. The gram formula mass of sodium chloride contains 6.02×10^{23} formula (NaCl) units.

Summary

6.02×10^{23} atoms of sodium weigh 22.99 grams – 22.99 g/mole
6.02×10^{23} molecules of water weigh 18.00 grams – 18.00 g/mole
6.02×10^{23} formula units of sodium chloride weigh 58.5 grams – 58.5 g/mole

Experiment 11.2

Equipment and Supplies per Lab Station
sample of an element; sample of a compound
Class Supplies
stoppered graduated cylinder containing one mole of water
stoppered graduated cylinder containing one mole of glycerin
stoppered bottles of each of the following:
1 mole of iron filings, 1 mole of copper pellets
1 mole of silicon dioxide, 1 mole of table sugar

What to do

PART I

1. Locate two stoppered graduated cylinders on the supply table. One contains one mole of water, H_2O. The other contains one mole of glycerin, $CH_2OHCHOHCH_2OH$.
   Observe and record the volume and grams per mole of each in Data Chart 1.
   Calculate the density of both water and of glycerin. Record these values in Data Chart 1.

2. Locate four stoppered bottles on the supply table. They contain the following samples:
   Bottle #1 1 mole of iron (iron filings) – Fe
   Bottle #2 1 mole of copper (copper pellets) – Cu
   Bottle #3 1 mole of silicon dioxide (pure sand) – SiO_2
   Bottle #4 1 mole of sugar (table sugar) – $C_{12}H_{22}O_{11}$
   Record the volume of one mole of each as given on the label.
   Find the atomic mass of iron and the atomic mass of copper in the Periodic Table. Record these in terms of grams per mole.
   The grams per mole for silicon dioxide and for sugar are given.
   Calculate the density for each of the four substances and record. Complete the Data Chart.

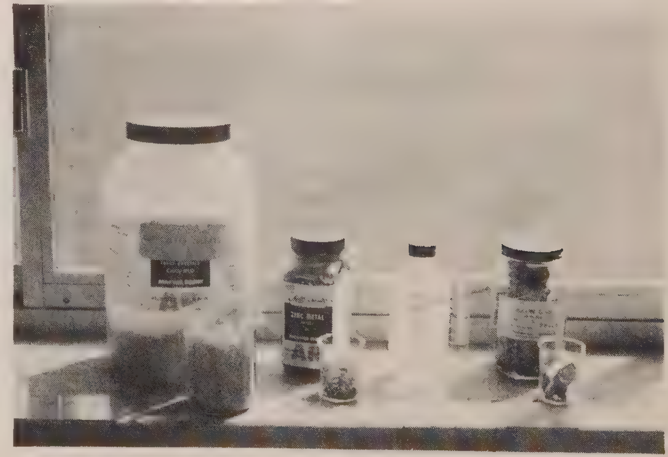
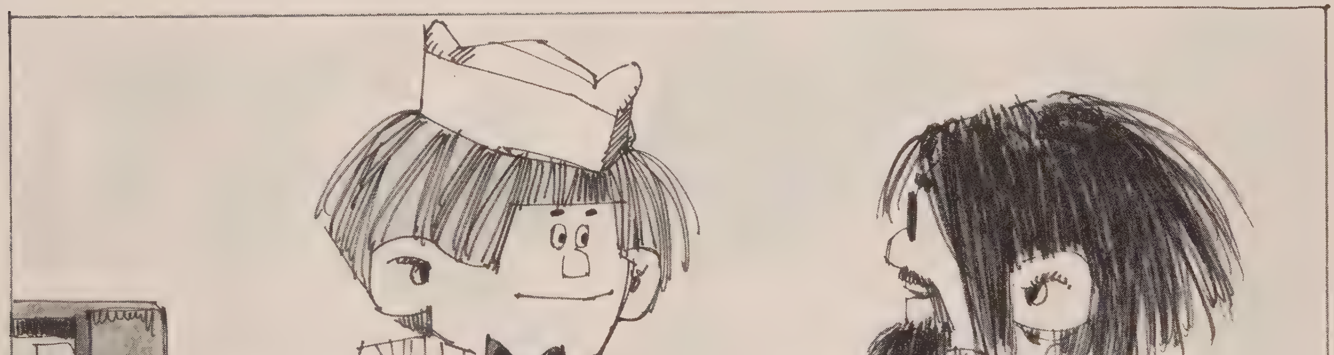


Fig. 53. The mole mass of each of the following has been placed on the laboratory shelf: copper(II) sulfate; zinc; water; and silicon. *(Clara Aich and George Senty)*

DATA CHART 1

Substance	Unit Particles Composing the Substance	Total Number of Unit Particles in One Mole of Substance	Mass in Grams of One Mole of Substance	Volume in ml of One Mole of Substance	Density at Room Conditions D = g/ml
1 mole of H_2O	molecules	6.02×10^{23}	18 g		
1 mole of $CH_2OHCHOHCH_2OH$	molecules	6.02×10^{23}	92.09 g		
1 mole of Fe	atoms				
1 mole of Cu	atoms	6.02×10^{23}			
1 mole of SiO_2	molecules		60.06 g		
1 mole of $C_{12}H_{22}O_{11}$	molecules	6.02×10^{23}	342 g		


Give me a mole of eggs, please. And charge it.

PART II

What to do

1. Obtain a sample of an element.
2. Weigh the sample. Record the mass of the sample in Data Chart 2.
3. Look up the atomic mass of the element in the Periodic Table. Record in terms of grams per mole in Data Chart 2.
4. Determine the number of moles in the sample. Record in Data Chart 2.

$$\frac{\text{mass of the sample (g)}}{\text{gram atomic mass (g/mole)}} = \text{number of moles in the sample}$$

5. Obtain a sample of a compound.
6. Weigh the sample. Record the mass of the sample in Data Chart 2.
7. Calculate the formula mass from the formula. Use the Periodic Table to find the atomic mass of each atom in the formula. Multiply this mass by the number of atoms of that element in the formula.

(a) Example: Na_2SO_4	No. of atoms		Atomic mass		Totals
atomic mass of Na = 22.99	2	×	22.99	=	45.98 g/mole
atomic mass of S = 32.10	1	×	32.10	=	32.10 g/mole
atomic mass of O = 16.00	4	×	16.00	=	64.00 g/mole
mass of 1 mole of Na_2SO_4	. . . .				142.08 g/mole

(b) Gram formula mass (g/mole) of your compound:

Formula: ________	No. of atoms		Atomic mass	Totals
atomic mass of ___ = ___	________	×	________	______ g/mole
atomic mass of ___ = ___	________	×	________	______ g/mole
atomic mass of ___ = ___	________	×	________	______ g/mole
mass of 1 mole of ________	. . . .			______ g/mole

Record the gram formula mass (g/mole) of your compound in Data Chart 2.

8. Determine the number of moles in your sample of the compound.

$$\frac{\text{mass of the sample (g)}}{\text{gram formula mass (g/mole)}} = \text{number of moles in the sample}$$

9. Repeat steps 5, 6, 7, and 8 with as many samples of different compounds as time permits.

DATA CHART 2

Substance		Mass of Sample (grams)	Formula Mass (grams/mole)	No. of Moles in Sample
Name	Formula			

Questions to answer

1. How do the volumes of water and glycerin compare? ______________________________
____

2. How do molar volumes of solids compare?
____
____

3. Compare the densities of the substances with which you worked. ______________________________
____
____

Extra for experts

How would you know when you had a mole of a substance for which you knew the formula?

____
____

(John H. Gerard from National Audubon Society)

GRAM FORMULA MASS—GRAMS PER MOLE

Calculate the gram formula mass, grams per mole, for each of the following compounds. Refer to the calculation of the grams per mole of sodium sulfate in Experiment 11.2, PART II.

1. Sodium bromide, NaBr

2. Nitric acid, HNO_3

3. Barium nitrate, $Ba(NO_3)_2$
   Note: *The subscript beside the parenthesis multiplies the number of the atoms inside the parenthesis. Here there are two atoms of nitrogen and six atoms of oxygen.*

4. Iron III sulfate, $Fe_2(SO_4)_3$

5. Water, H_2O

6. Sodium hydroxide, NaOH

7. Calcium carbonate, $CaCO_3$

8. Lithium iodide, LiI

9. Silver chloride, AgCl

10. Glucose, $C_6H_{12}O_6$

Exercise

HOW MUCH OF THE ELEMENT IS COMBINED

The percent of an element present in a compound can be calculated as in the following example.

Example:

1. The formula mass of calcium chloride, $CaCl_2$.

	No. of atoms	Atomic mass		Totals
atomic mass of calcium = 40.0	1	× 40.0	=	40.0 g/mole
atomic mass of chlorine = 35.5	2	× 35.5	=	71.0 g/mole
calcium chloride, $CaCl_2$				111.0 g/mole

2. The percent of calcium in calcium chloride, $CaCl_2$.

$$\frac{40.0 \text{ g/mole of calcium in } CaCl_2}{111 \text{ g/mole of } CaCl_2} \times 100 = 36\%$$

What to do

Calculate the following percents of elements in compounds.

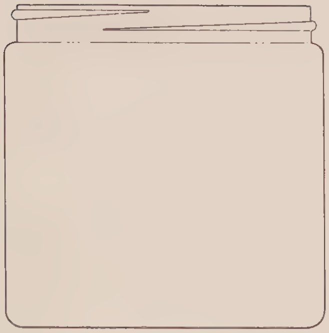
1. The percent of sodium in sodium oxide, Na_2O.

2. The percent of copper in copper nitrate, $Cu(NO_3)_2$.

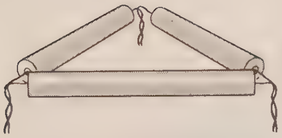
3. The percent of hydrogen in sulfuric acid, H_2SO_4.

HOW TO

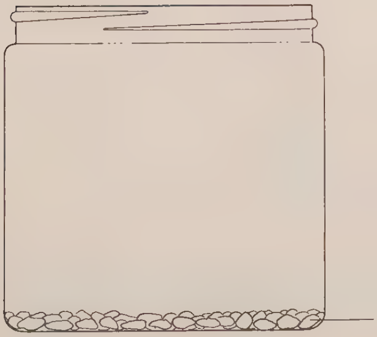
HOW TO MAKE A DESSICATOR



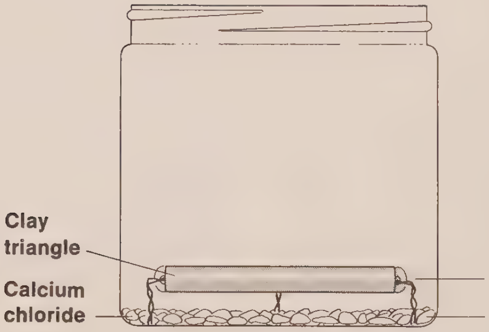
1. Obtain a wide-mouthed jar with a tight sealing lid, like a peanut butter jar.



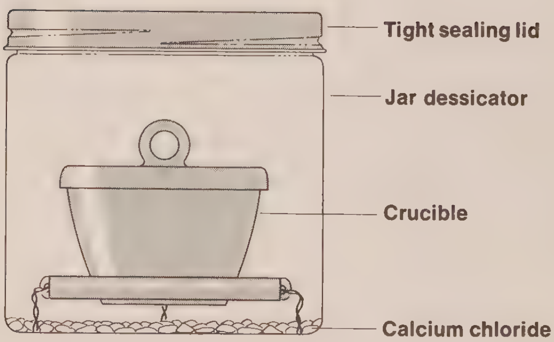
2. Bend the prongs on a clay triangle to make legs of even length.



3. Cover the bottom of the jar with a thin layer of calcium chloride, $CaCl_2$.



4. Fit the clay triangle holder snugly into the jar.



5. If a test tube is to be placed in the dessicator place a wire screen over the clay triangle to prevent the test tube from resting on the calcium chloride.

Burnt Sugar

Have you ever burned something with sugar in it? It may have been candy, cake, or frosting. It gets all hard and black. What a mess!

What happens to sugar when it "burns"? What is it made of anyway?

The chemist knows table sugar by its formula, $C_{12}H_{22}O_{11}$. This means that in one mole of sugar, 12 moles of carbon atoms are combined with 22 moles of hydrogen atoms and eleven moles of oxygen atoms. Sugar breaks up when it is heated. The hydrogen and oxygen go away as water vapor and leave the black carbon.

In this experiment you will burn a sample of sugar and find the percent of water lost.

Experiment 11.3

Equipment and Supplies per Lab Station

test tube (may be discarded after the experiment)
balance, centigram; burner; buret clamp
dessicator*, student made; ring stand
½ teaspoonful of sucrose, $C_{12}H_{22}O_{11}$

*If the student makes his own dessicator, it must be tall enough for a test tube to be placed into it. The test tube takes the place of the crucible shown in it on page *4*/**49**.
Materials needed for student-made dessicator:
jar with tight-fitting lid
clay triangle; anhydrous calcium chloride; wire screen

Coffee please. Milk and two lumps of $C_{12}H_{22}O_{11}$.

What to do

1. Attach a clean dry test tube to a ring stand with a clamp.
2. Heat the empty test tube gently about three minutes. Cool it until it is just barely warm. Cool in a dessicator, if possible. See "How to Make a Dessicator" on pg. *4*/49.
3. Weigh the cool test tube to the nearest 0.1 g. Record the mass in the Data Chart.

4. Transfer about one-half teaspoon of table sugar to the test tube. Weigh the test tube and sugar to the nearest 0.1 g. Record the mass in the Data Chart.

5. Heat the sugar in the test tube *gently* for about two minutes.

   (a) What appears in the cooler part of the test tube near the top? ______________________

   ____

   (b) What changes do you see in the sugar?

   ____

   ____

6. Increase the heat. Continue heating until there is no more change in color and the inside of the test tube is dry. Describe what is left in the test tube. ______________________

   ____

7. Cool the test tube and contents until they are barely warm. Use the dessicator, if possible. Weigh the cool test tube to the nearest 0.1 g. Record the mass in the Data Chart.

8. Repeat steps 6 and 7 until there is no change in the mass of the test tube. Clean the test tube as well as you can. Your teacher will give you advice.

DATA CHART

Mass of the dry test tube	____g
Mass of the test tube and sugar before heating	____g
Mass of the test tube and char after heating	____g

CALCULATIONS

1. Mass of the sugar before heating ____g
   (Subtract the mass of the test tube from the mass of the test tube plus the sugar.)

2. Mass of the water lost ____g
   (Subtract the mass of the test tube plus the char *after* heating from the mass of the test tube plus the sugar *before* heating.)

3. Percent of water lost ____%

$$\frac{\textit{mass of water lost}}{\text{mass of sugar}} \times 100 = \_\_\_\_\% \text{ of water lost}$$

HOW ACCURATE ARE YOU?

1. Calculate the molecular mass of water, H_2O.
hydrogen = (1 g per mole) × (2 atoms per mole) = 2 g
oxygen = (16 g per mole) × (1 atom per mole) = 16 g
H_2O .... = 18 g/mole

2. Calculate the molecular mass of table sugar, $C_{12}H_{22}O_{11}$.

	g/mole	No. of atoms	Totals
atomic mass of carbon	= ______ ×	______ =	______ g/mole
atomic mass of hydrogen	= ______ ×	______ =	______ g/mole
atomic mass of oxygen	= ______ ×	______ =	______ g/mole
formula mass of $C_{12}H_{22}O_{11}$	....	=	______ g/mole

3. Calculate the grams of water in one mole of table sugar.
(18 g/mole of water) × (11 moles of water in

table sugar) = ____ g

4. Calculate the percent of water in the gram molecular mass of table sugar, $C_{12}H_{22}O_{11}$.

$$\frac{\text{gram molecular mass of 11 } H_2O}{\text{gram molecular mass of } C_{12}H_{22}O_{11}} \times 100 =$$

____% of water

5. How well does the percent of water lost in your experiment agree with the percent of water in

one mole of sugar, $C_{12}H_{22}O_{11}$? ______________________

____

How do you account for any difference? ________

____

____

Questions to answer

1. Why do you heat the empty test tube before

doing the test? ______________________

____

2. Why must you cool the test tube before

weighing it? ______________________

____

3. Why is it better to cool the test tube in a

dessicator rather than out in the room? ________

____

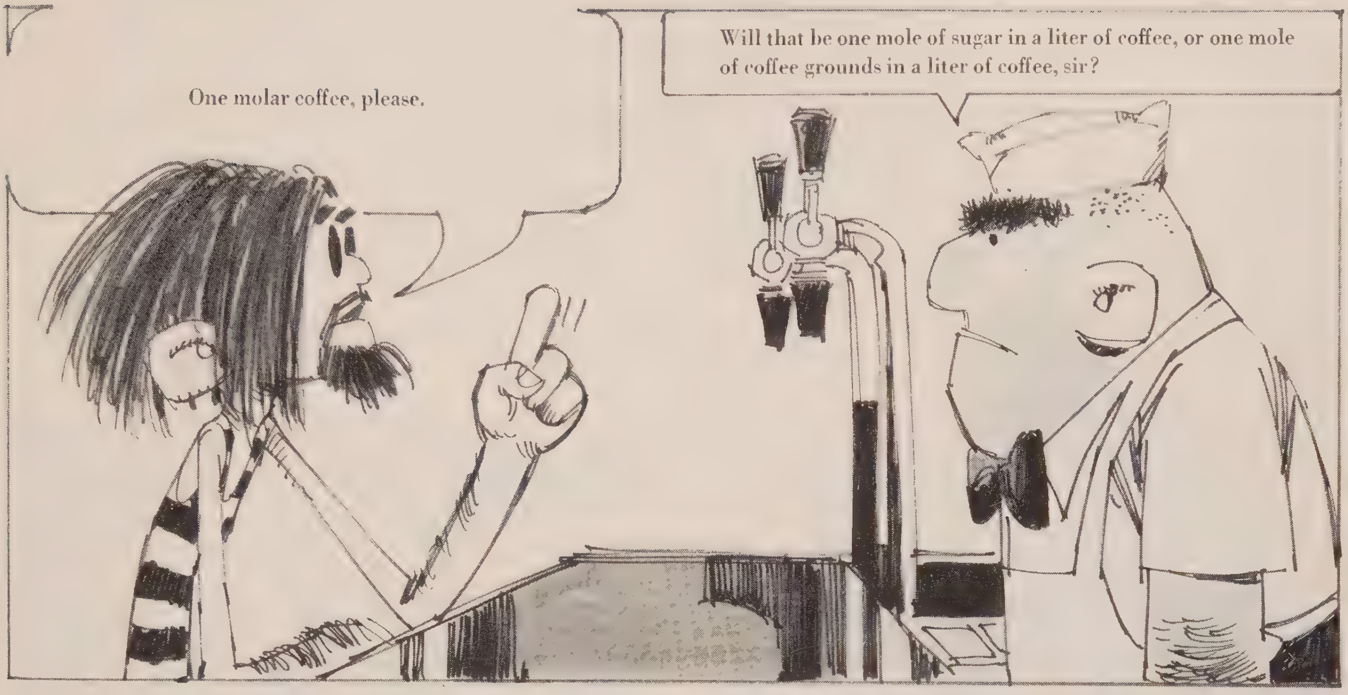
____

4. How would you describe what happens to sugar

when it is heated? ______________________

____

____


One molar coffee, please.
Will that be one mole of sugar in a liter of coffee, or one mole of coffee grounds in a liter of coffee, sir?

MAKING MOLAR SOLUTIONS FROM SOLIDS

Most of the solutions you use in the laboratory are molar solutions. Their labels include such symbols as $2M$, M, $0.2\ M$, $0.5\ M$, and $0.1\ M$. The capital M stands for MOLAR. The numbers show the concentration in moles of solute per liter of solution.

A ONE MOLAR SOLUTION contains one mole of solute (the dissolved substance) in one liter of solution. Other molarities are either multiples of the mole or decimal fractions of the mole quantity.

Your teacher will prepare a $1\ M$ solution of sodium hydroxide, NaOH.

You will prepare a $0.1\ M$ solution of sodium hydroxide, NaOH for use in Experiments 12.2 and 12.4.

Experiment 12.1

Equipment and Supplies per Lab Station
volumetric flask, 1 liter—for demonstration
volumetric flask, 1 liter—for student
or graduated cylinder, 1 liter
beaker—150 ml; stirring rod; balance, centigram
flask, 1 liter—if graduated cylinder is used by student
(2) stoppers to fit the flasks; distilled water
50 g sodium hydroxide pellets:
  40 g for the demonstration
  4 g for the student

What to do

1. TEACHER DEMONSTRATION

(a) Weigh a small beaker. Record the mass. Add the mass of one mole of sodium hydroxide, NaOH, to the beams of the balance. Your teacher will tell you what these masses are after you have made your own calculations.

DATA

Mass of the beaker ____________g

Mass of one mole of NaOH ____________g

Mass of the beaker plus one mole of NaOH ____________g

(b) Add NaOH pellets until the mass on the balance is just balanced. See Fig. 54.

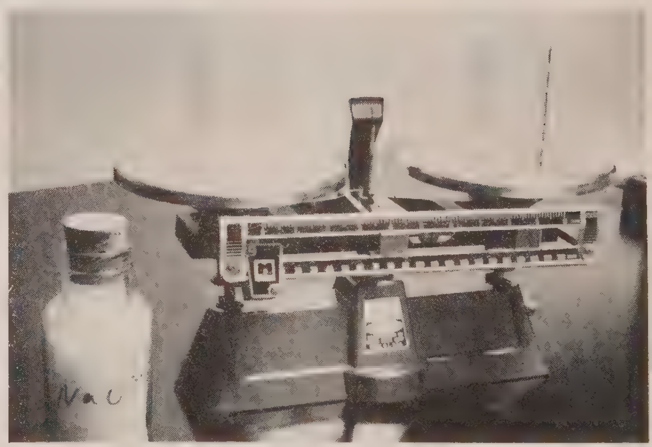


Fig. 54. A beaker of the same size may be placed on the right pan of the balance to counterbalance the mass of the beaker. Such a counterbalance is known as a tare. *(Clara Aich and George Senty)*

(c) Remove the beaker from the balance. Add about 20 ml of distilled water to the beaker. Stir to dissolve the NaOH. Let the undissolved NaOH settle. See Fig. 55.

(d) Pour the clear solution down a stirring rod into a liter volumetric flask (or liter graduated cylinder). See Fig. 56.

(e) Again add a small amount of distilled water to the beaker. Stir. Pour the solution down the stirring rod into the volumetric flask. Repeat until all of the NaOH is dissolved. Be careful not to use too much distilled water. You should not have more than 995 ml when the NaOH is all dissolved.

(f) Last, add distilled water to the volumetric flask with a medicine dropper until you have just 1000 ml of solution. See Fig. 57. Stopper tightly. Mix thoroughly by inverting many times. See Fig. 58.

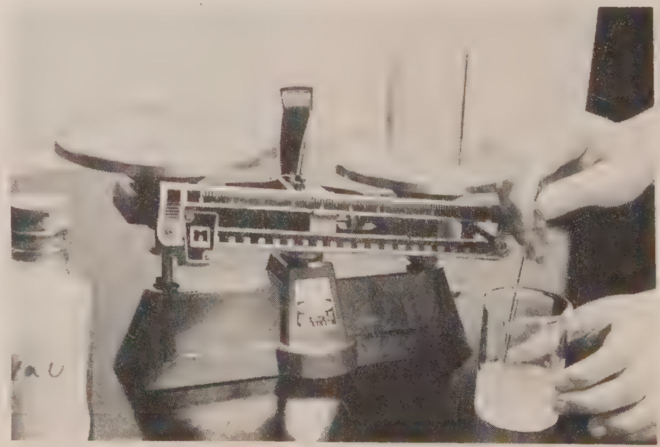


Fig. 55. *(Clara Aich and George Senty)*

2. Your preparation of 0.1 *M* NaOH
   (a) By dilution
   Obtain 100 ml of the 1 *M* solution of NaOH. Transfer this to a one liter graduated cylinder or volumetric flask. Add distilled water until you have one liter, 1000 ml of solution. Add the last water drop by drop so that you do not add too much. If you have used a one liter graduated cylinder, transfer the solution to a one liter flask. Stopper the flask tightly. Mix the solution thoroughly by turning the flask upside down and back many times. Do not get any solution on your hands. If you do, wash with water at once.

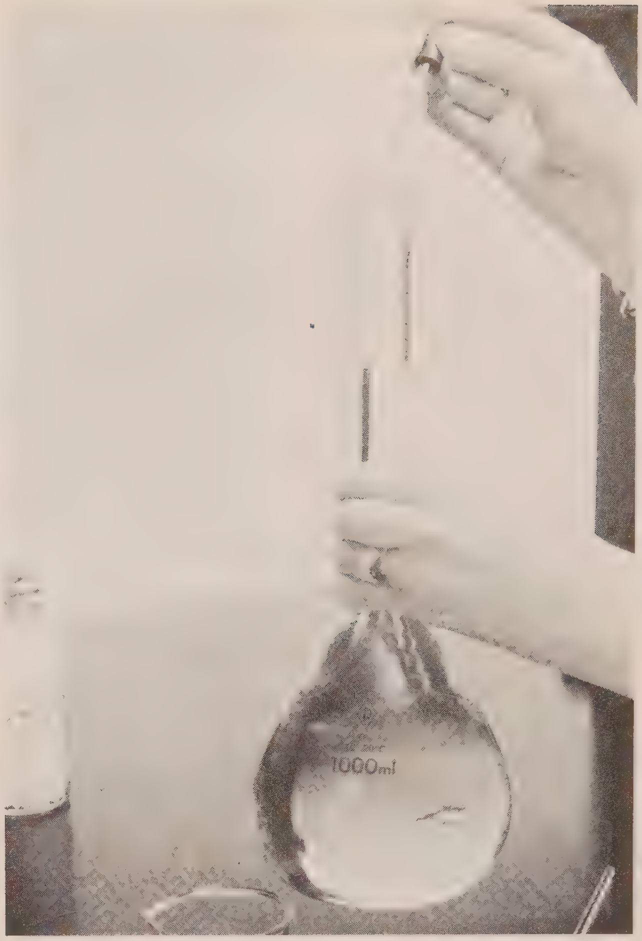


Fig. 57. *(Clara Aich and George Senty)*

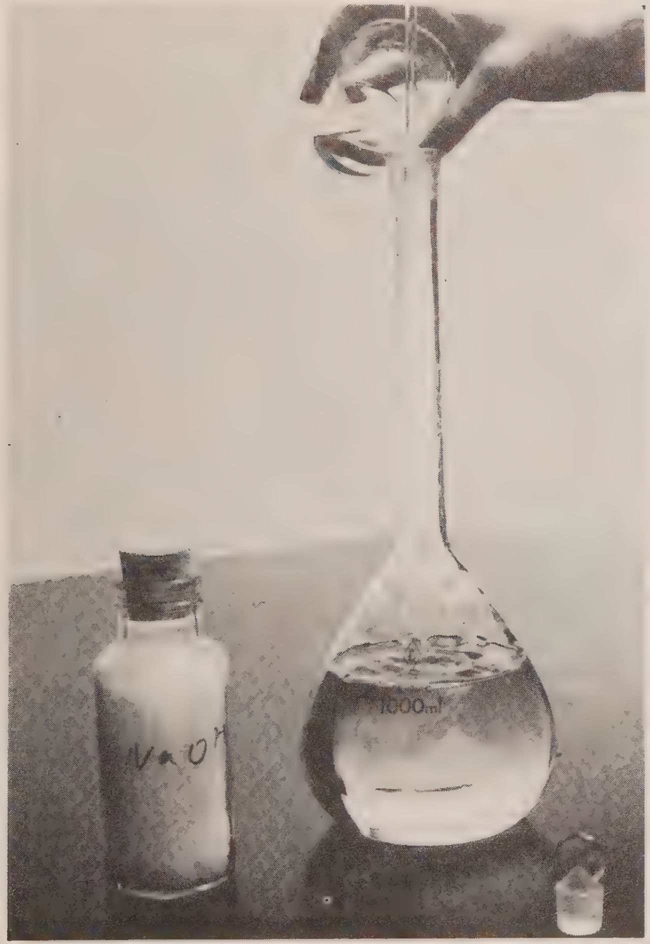


Fig. 56. *(Clara Aich and George Senty)*

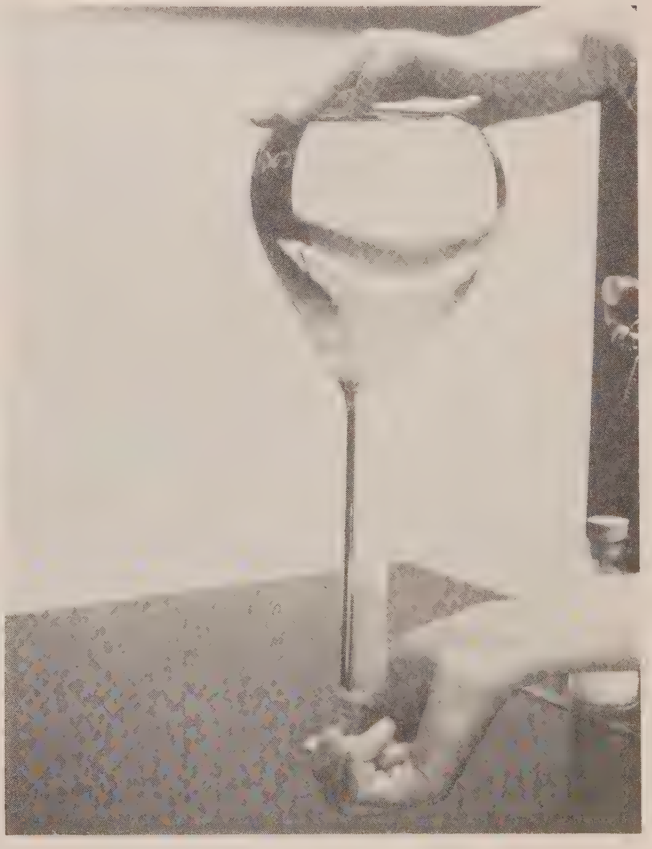


Fig. 58. *(Clara Aich and George Senty)*

(b) By calculation and weighing
Calculate the formula mass of one mole of NaOH by the method used in the exercise in Chapter 11.
Gram formula mass of sodium hydroxide,

NaOH = ____________ g/mole
Divide the grams per mole of sodium hydroxide by ten to find the number of grams needed for a 0.1 *M* solution.

$$\frac{\text{grams per mole of NaOH}}{10} =$$

__________ grams per 0.1 mole of NaOH
Weigh, dissolve, and make one liter of solution from this number of grams just as your teacher did from the quantity for one liter of a one molar solution.

(c) Save your solution for use in Experiments 12.2 and 12.4.

Questions to answer

1. What difference would incorrect calculations or incorrect weighing make? ____________________

____

____

2. What difference would spilling some of the solution in transferring make? ________________

____

3. Why does it matter whether or not all of the NaOH is washed out of the beaker in which it is weighed and dissolved? ________________

____

____

4. What difference would it make if you added distilled water beyond the 1000 ml mark?

____

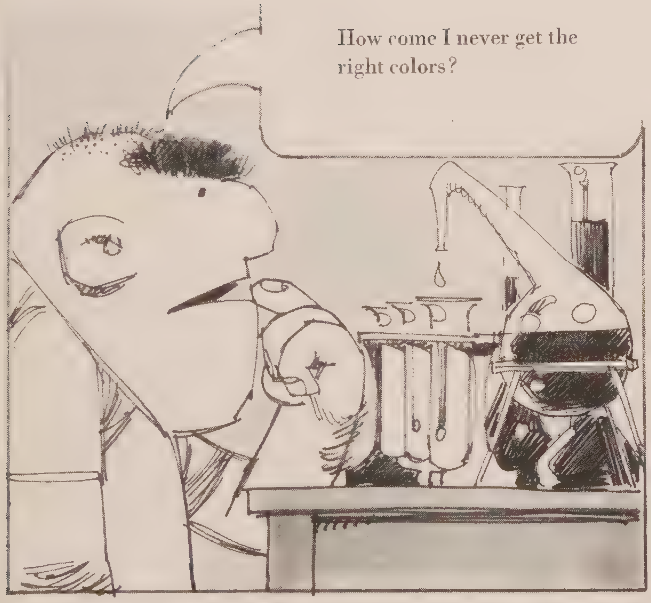
____

____

DELVING FURTHER INTO pH

In Chapter 8 you learned that a base can neutralize an acid, that is, a base can cancel the acid properties of an acid. At the neutral point all of the hydrogen ion (H^+) of the acid and all of the hydroxyl ion (OH^-) of the base have reacted to form water. Similarly, an acid can neutralize a base. This process may be represented on the pH scale.

← H^+ INCREASING — H^+ decreasing →
neutral
↓
pH 0 1 2 3 4 5 6 7 8 9 10 11 12 13 14
$H^+ = OH^-$
← OH^- decreasing — ——— OH^- INCREASING →
Acid pH readings End point Basic pH readings
in a titration


How come I never get the right colors?

Experiment 12.2

Equipment and Supplies per Lab Station
(22) test tubes, 18 × 150 mm; graduated cylinder, 25 ml or 50 ml
(4) test tube racks, 6 places each; distilled water
vial Universal Indicator papers with color key
dropper bottle, 30 ml, of phenolphthalein
dropper bottle, 30 ml, of methyl orange
60 ml 0.1 *M* hydrochloric acid; 60 ml 0.1 *M* sodium hydroxide

What to do

1. Prepare eleven test tubes with the quantities of 0.1 *M* hydrochloric acid and 0.1 *M* sodium hydroxide shown in the first columns of the chart. Use the 0.1 *M* sodium hydroxide solution you prepared in Experiment 12.1. Mix each test tube well.
2. Transfer one drop from each test tube to a separate strip of Universal Indicator paper. Compare the color formed with the color chart on the vial. Record the color and the pH in the table.
3. Pour half of the solution from each test tube into another test tube. You now have two identical sets of solutions.
4. Add to each test tube in one set two drops of phenolphthalein indicator. Record the color for each test tube in the chart.
5. Add two drops of methyl orange to each test tube in the second set. Record color changes.

Indicator colors

Indicator	*Acid*	*End point*	*Base*
Phenolphthalein	colorless	pink	red
Methyl orange	red	orange	yellow
Universal Indicator	consult color chart		

COMPARISON OF INDICATORS

Test Tube	0.1M HCl	0.1M NaOH	Universal Indicator		Phenolphthalein	Methyl orange
			Color	pH		
#1	10 ml	0 ml				
#2	9 ml	1 ml				
#3	8 ml	2 ml				
#4	7 ml	3 ml				
#5	6 ml	4 ml				
#6	5 ml	5 ml				
#7	4 ml	6 ml				
#8	3 ml	7 ml				
#9	2 ml	8 ml				
#10	1 ml	9 ml				
#11	0 ml	10 ml				

Questions to answer

1. At what pH did you find the change point of each indicator?

   Universal Indicator pH ______

   Phenophthalein pH ______

   Methyl Orange pH ______

2. (a) The total pH range is: from pH ______ to pH ______

   (b) The acid pH range is: from pH ______ to pH ______

   (c) The basic pH range is: from pH ______ to pH ______

Extra for experts

1. What makes the bubbles in a soft drink? ______

   ____

2. Would you expect the soft drink solution to be acidic or basic? Explain. ______________

   ____

3. How would you test the pH of a soft drink? ______

   ____

   ____

A LITTLE SUPER-SUPER DISHWASHING
HOW TO USE BURETS

Burets must be super clean. They are used to measure small volumes of liquids when accuracy is essential.

What does super clean mean?

How do you use burets?

Experiment 12.3

Equipment and Supplies per Lab Station
beaker—250 ml; buret—50 ml; Erlenmeyer flask—250 ml buret brush; buret stand with double pinch clamp for two burets; or ring stand with screw type clamps; detergent solution; distilled water

What to do

1. Super cleaning the buret
    - (a) Place about 10 ml of a detergent solution in the buret. Scrub the inside surface with a buret brush.
    - (b) Rinse well, first with tap water and finally with distilled water. Pour the water into the buret from a beaker or flask.
    - (c) Drain the buret in a vertical position. If any droplets of water cling to the inside surface, the buret is *not* clean. Rewash, rerinse, and drain the buret until it is clean.
2. Preparing the buret for use
    - (a) Place 10–15 ml of a solution to be measured in the clean buret. *Note:* Distilled water may be used for practice instead of a solution used in an experiment.
    - (b) Allow 3–5 ml to flow through the delivery tip. See Fig. 59.
    - (c) Turn the buret to the horizontal position. Rotate it to spread the solution evenly over the inner surface. Pour the solution out of the top. Make sure that the entire inner surface has been made wet by the solution.
    - (d) Fill the buret to the top with the solution. Force any air bubbles out of the delivery tip. See Fig. 59.
    - (e) Let the solution flow from the tip until the bottom of the meniscus in the buret rests at zero or below. Record volume.

      Note 1: A drop may hang on the tip before you start a real test. If so, discard it by touching it to a beaker for waste liquid as shown in Fig. 59.

      Note 2: A drop formed during a test must be caught by touching it to the inside of the flask being used for the test. Rinse it into the flask with a little distilled water.
3. Reading the volume
    - (a) Be sure to have your eye on a level with the bottom of the meniscus. Read the volume at the bottom of the curve carefully to the nearest 0.1 ml. Record each volume read during a test.
    - (b) Use an absolute minimum of 10 ml of the

solution to be tested in each test made. Less than this increases the error in the measurement.

4. Burets after use
   (a) *Acid buret:* Drain out any remaining acid and discard. Rinse the buret several times with tap water. Finally, rinse at least three times with distilled water. Drain dry in a vertical position. Store in a clean safe place.

   (b) *Base buret:* Drain out and discard any remaining base solution. Rinse the buret several times with tap water. Now rinse with an acid solution. Rinse again several times with tap water. Finally, rinse three times with distilled water. Drain dry in a vertical position. Store in a clean safe place.

   Note: *Solutions of bases tend to react with glass. The acid rinse removes any base which might react with the glass. The buret used to measure a solution of a base must be cleaned immediately after use.*

Questions to answer

1. Why must no droplets of liquid cling to the inside of the buret? ____________________

   ____

   ____

   ____

2. Why must air bubbles be forced out of the delivery tip? ____________________

   ____

3. (a) Why must a drop of liquid formed on the delivery tip before the test be removed and discarded? ____________________

   ____

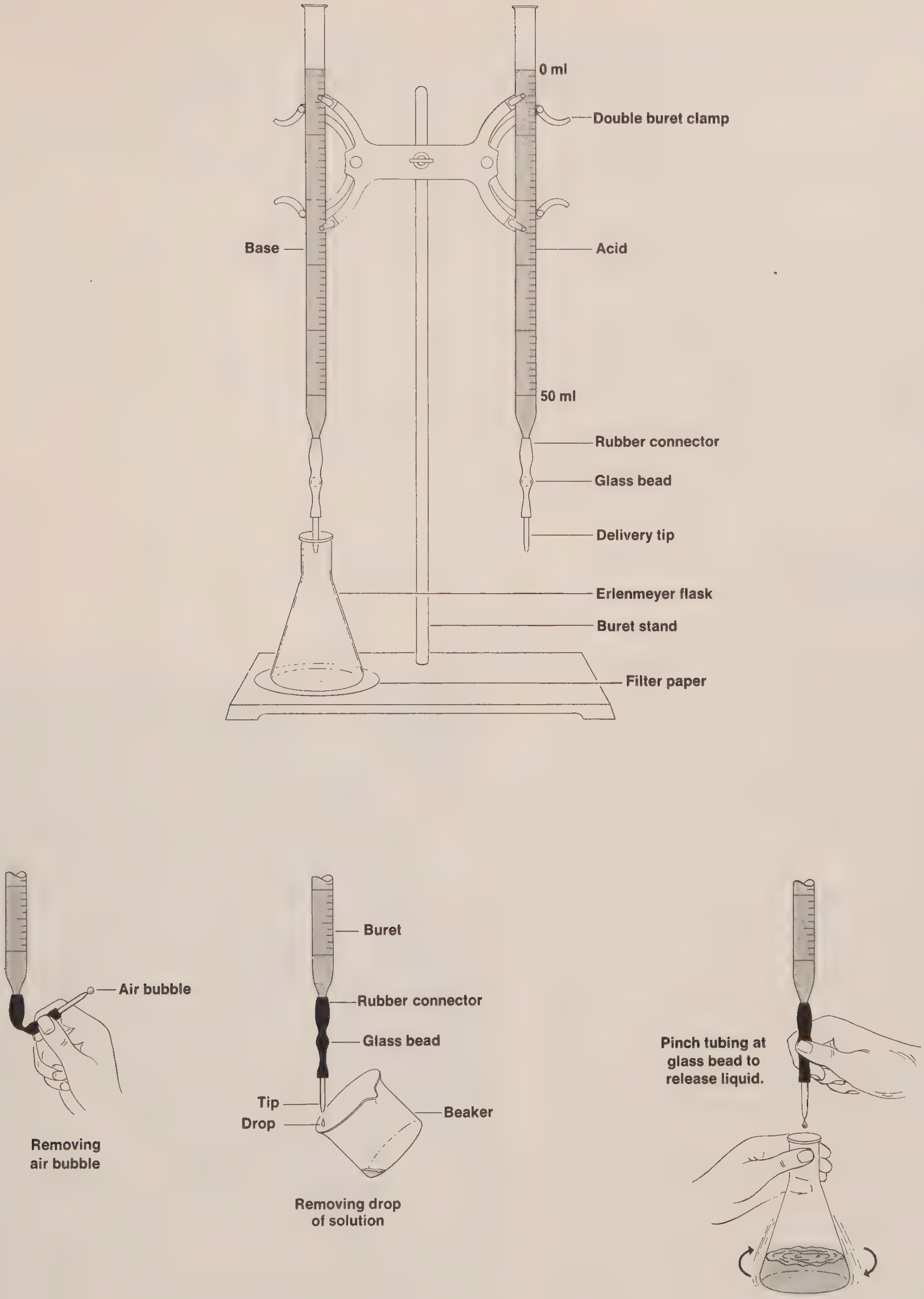
   (b) Why must a drop of liquid formed during a test be rinsed into the flask in which the test is being done? ____________________

   ____

   ____

4. Why should the solution of a base never be left standing in a buret? ____________________

   ____



Removing air bubble

Removing drop of solution

Fig. 59. Titration apparatus. The burets have connectors made of rubber.

HOW SOUR IS VINEGAR?

The sourness of vinegar is due to acetic acid. Perhaps a titration will show how sour vinegar is.

TITRATION is the addition of measured amounts of a solution of known concentration to a measured volume of a solution of unknown concentration.

In acid-base titrations, indicators which change color are added to show when the reaction is complete.

The end-point of a titration is reached when the combination of acid and base is neutral to the indicator. It is neutral to the indicator when one drop of either acid or base causes a change in the color of the indicator.

Experiment 12.4

Equipment and Supplies per Lab Station

(2) burets – 50 ml; (2) flasks – 250 ml; (2) beakers – 250 ml
buret support; 1 dropper bottle of phenolphthalein
distilled water; 100 ml 0.1 *M* sodium hydroxide solution
10 ml grocery store colorless distilled vinegar
graduated cylinder – 10 ml or 25 ml
graduated cylinder – 100 ml or 50 ml

What to do

1. Clamp two clean 50-ml burets to a support as illustrated in Fig. 60.
2. Rinse one of the burets with the 0.1 *M* sodium hydroxide solution you prepared in Experiment 12.1. Empty the buret.
3. Fill the buret with the 0.1 *M* sodium hydroxide from Experiment 12.1 to about the 50 ml mark. Label the buret "base". Force air bubbles out of the delivery tip as you did in Experiment 12.3. Read the volume of the 0.1 *M* sodium hydroxide to the nearest 0.1 ml. Record this reading in your Data Chart.
4. Transfer exactly 10 ml of vinegar to a graduated cylinder. Pour this vinegar into 85 ml of distilled water in a 100 ml graduated cylinder. Add distilled water with a medicine dropper up to the 100 ml mark. Transfer this solution to a 250-ml Erlenmeyer flask. Stopper the flask and mix thoroughly. This solution will be nearly the same strength as your 0.1 *M* solution of sodium hydroxide.

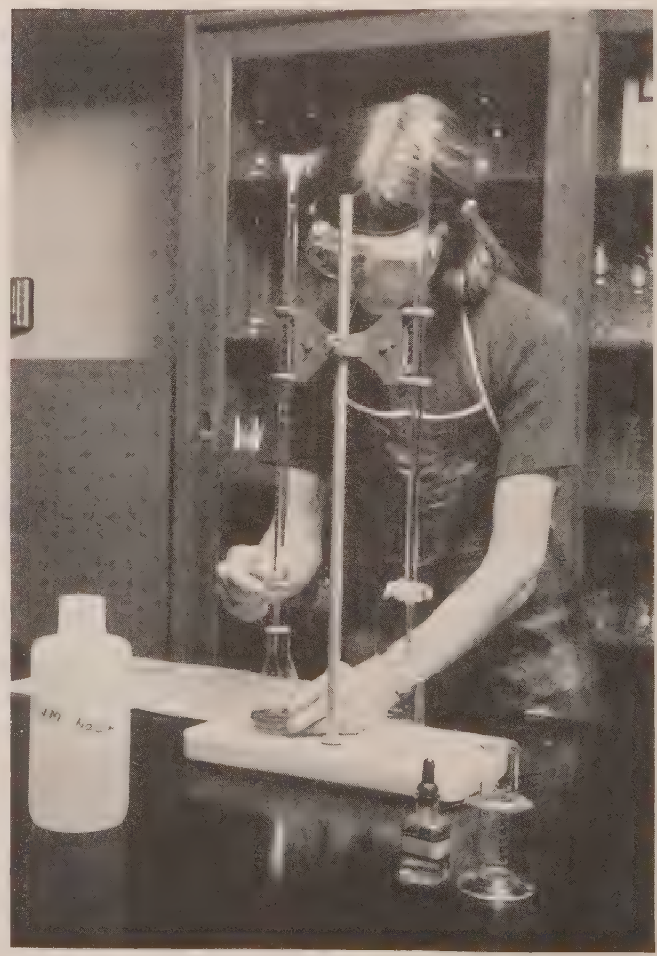


Fig. 60. Titration apparatus. The burets have glass stopcocks. *(Clara Aich and George Senty)*

5. Rinse the second buret with the vinegar solution. Empty the buret. Fill the buret to about 50 ml with the vinegar solution. Label the buret "acid". Force air bubbles out of the delivery tip. Read the volume of the vinegar solution to the nearest 0.1 ml. Record this volume in your Data Chart.

6. Titration
   (a) Measure 10 ml, as exactly as possible, of the vinegar solution into a clean 250-ml Erlenmeyer flask. Read the buret. Record the new volume in your Data Chart.

   (b) Add two drops of phenophthalein indicator. No color appears. Add 15–20 ml of distilled water if your teacher directs you to.

   (c) Add $0.1M$ sodium hydroxide solution slowly to the flask, keeping up a swirling motion of the flask, until one drop makes a pink color appear.

   (d) Add drops of vinegar solution until the color just disappears. Record the volume now of the vinegar solution in your Data Chart.

   (e) Add $0.1M$ sodium hydroxide drop by drop until the color again just appears. Record the final volume of $0.1M$ sodium hydroxide in your Data Chart.
   Note: *The* end-point *of this titration is the point where one drop of the 0.1*M *sodium hydroxide causes the red color just to appear.*

DATA CHART

Starting volume of 0.1 *M* sodium hydroxide	______ ml
Final volume of 0.1 *M* sodium hydroxide	______ ml
Starting volume of vinegar solution	______ ml
Volume after 10 ml of vinegar solution measured	______ ml
Final volume of the vinegar solution	______ ml

CALCULATIONS

1. Final volume of 0.1 *M* sodium hydroxide ______ ml
   minus Starting volume of 0.1 *M* sodium hydroxide ______ ml
   equals Volume of 0.1 *M* sodium hydroxide used ______ ml
2. Final volume of vinegar solution ______ ml
   minus Starting volume of vinegar solution ______ ml
   equals Volume of vinegar used ______ ml
3. What is the molarity of the vinegar solution?
   molarity of vinegar soln.

$$= 0.1\ M\ \text{(sodium hydroxide)} \times \frac{\text{ml of sodium hydroxide}}{\text{ml of vinegar solution}}$$

$$= \underline{\qquad}\ \text{M vinegar solution (due to acetic acid)}$$

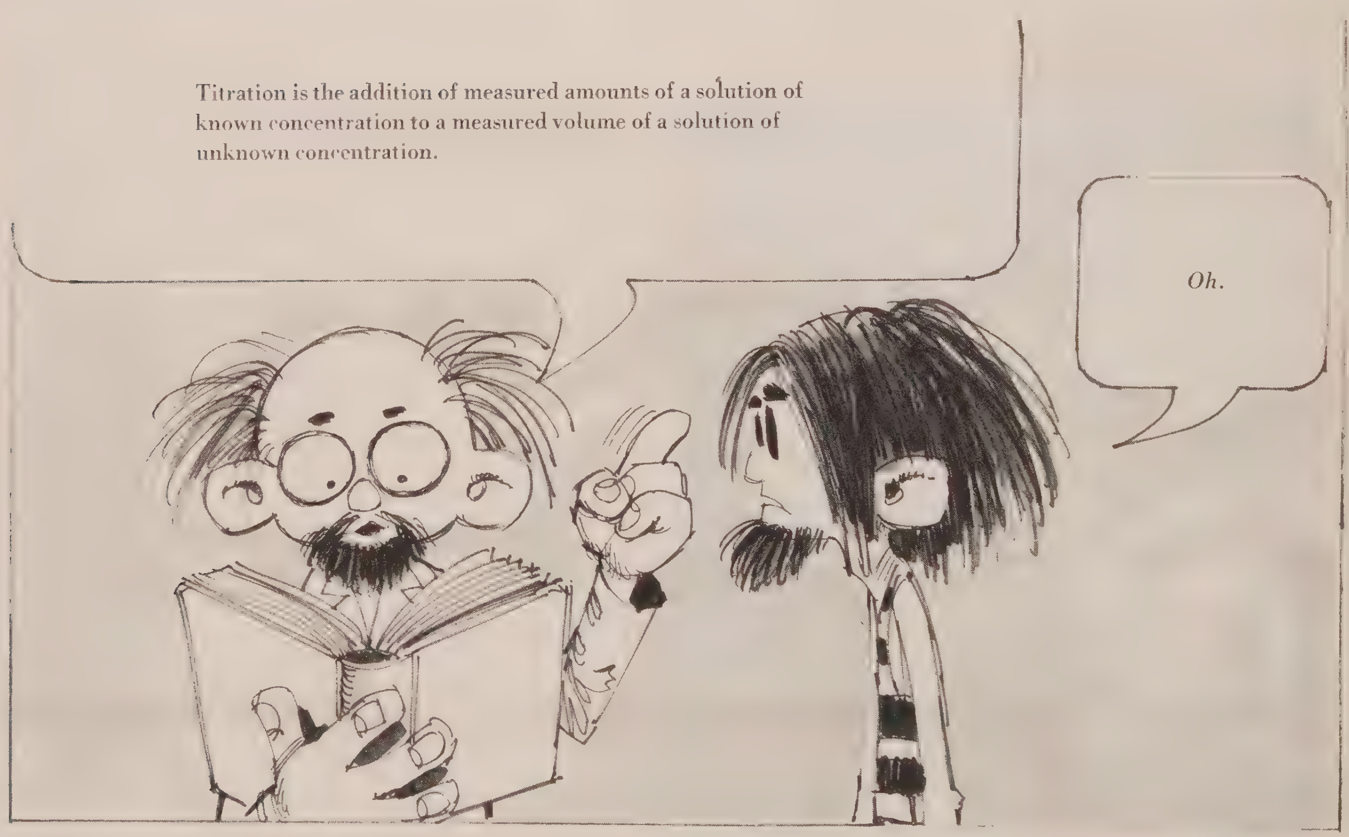
4. How sour is the vinegar?
   Diluted vinegar (from step 3 above) ______ M
   Undiluted vinegar (10 times the concentration) ______ M
   A $1M$ solution of acetic acid contains __60__ g/liter
   (Acetic acid, CH_3COOH, = 60 g/mole)

   Undiluted vinegar, ______ *M* acetic acid × 60 g/mole = ______ g/liter
   Assume that one liter of vinegar has a mass of 1000 grams.

$$\text{Then: } \frac{\text{grams/1000 ml of acetic acid in vinegar}}{\text{1000 ml}} \times 100 = \underline{\qquad}\%\ \text{acetic acid}$$

Example from an experiment like this one

1. Data: 10.3 ml of 0.1 M sodium hydroxide neutralized
   12.5 ml of diluted vinegar
2. Molarity of the vinegar
   (a) molarity of diluted vinegar
   $$= 0.1\ M\ \text{NaOH} \times \frac{10.3\ \text{ml of}\ 0.1\ M\ \text{NaOH}}{12.5\ \text{ml of diluted vinegar}}$$
   $$= 0.0824\ M\ \text{acetic acid (in diluted vinegar)}$$
   (b) Molarity of the undiluted vinegar (vinegar was diluted 10 times)
   $0.0824\ M \times 10 = 0.824\ M$ acetic acid (in undiluted vinegar)
3. The percent of acetic acid in vinegar
   (a) The grams of acetic acid in 0.824 M acetic acid (acetic acid, $CH_3COOH = 60$ g/mole)
   $0.824\ M \times 60\ \text{g/mole} = 49.4$ g/liter of acetic acid in vinegar
   (b) The percent of acetic acid in vinegar
   $$\frac{49.4\ \text{g/1000 ml of acetic acid}}{1000\ \text{ml}} \times 100 = 4.94\%$$
   acetic acid in vinegar


Titration is the addition of measured amounts of a solution of known concentration to a measured volume of a solution of unknown concentration.
Oh.

Questions to answer

1. Grocery store vinegar is usually about 3% to 5% acetic acid. How well does the percent of acetic acid in your sample of vinegar agree with this 3% to 5%? ____________________
____
2. How would you account for any difference? ____________________
____
____
____

Laboratory Experiments in

ACTION CHEMISTRY

UNIT V

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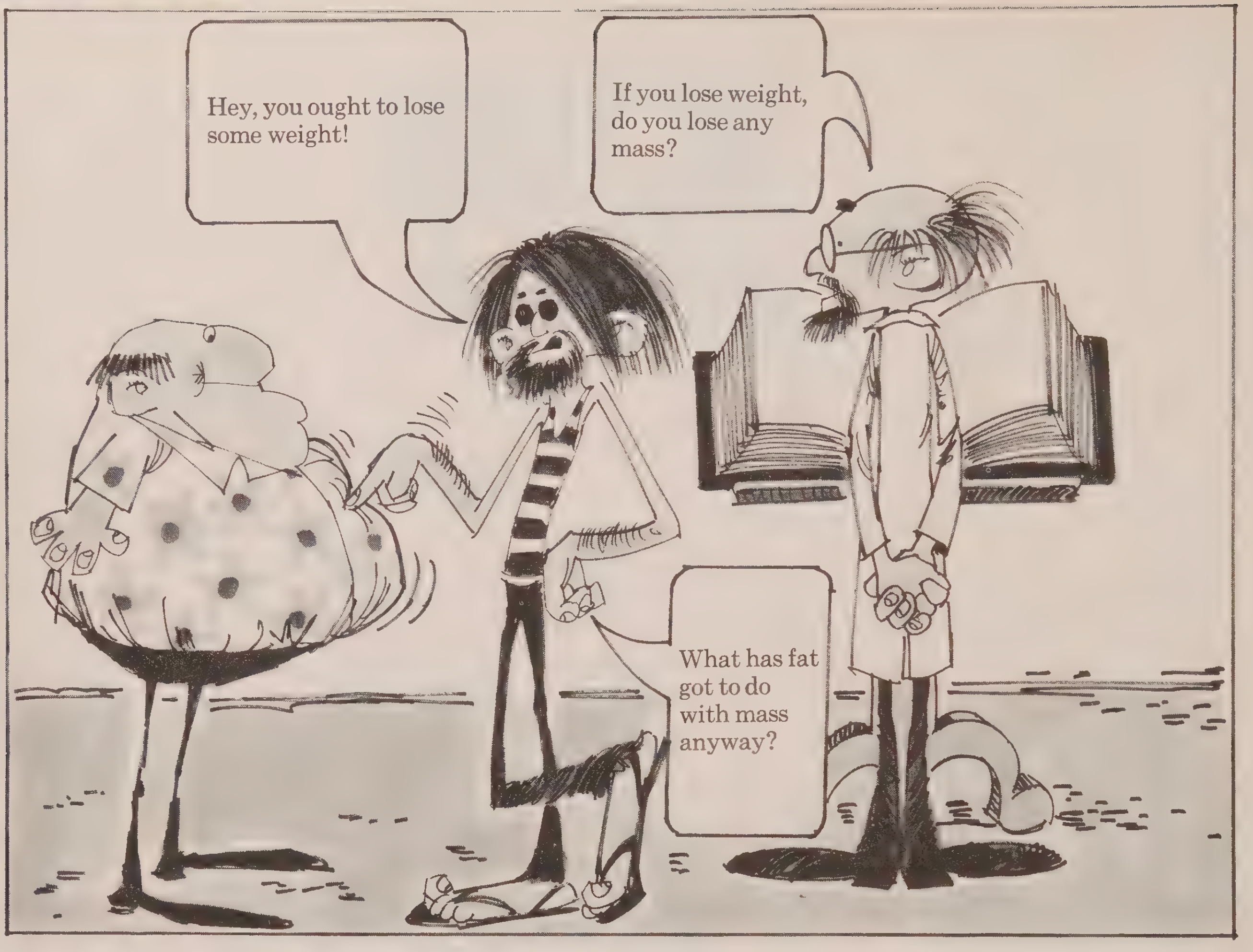

Hey, you ought to lose some weight!
If you lose weight, do you lose any mass?
What has fat got to do with mass anyway?

Conservation

Suppose you had two chemicals that will react if put together. If you measured the mass of the chemicals *before* they react and then measured the mass of the mixture *after* the reaction, how would the masses compare? Would the masses be different? Or would they be the same? If they are the same, mass is said to have been conserved. In this experiment, you will put two chemicals together. This will cause a chemical reaction.

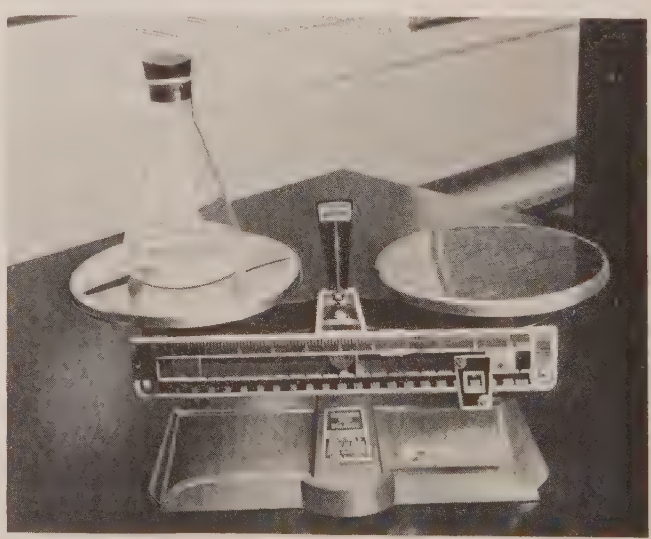
Using the information from this experiment, you should try to answer the question: How do the before and after masses compare?

Experiment 13.1

Equipment and Supplies per Lab Station

beaker—250 ml, Erlenmeyer flask—250 ml
graduated cylinder—25 ml or 50 ml
test tube which fits inside the flask (13 mm × 100 mm)
(2) test tubes (18 mm × 150 mm); (6) marble chips
10 ml of 0.1*M* solutions of:
barium chloride, $BaCl_2$; sodium carbonate, Na_2CO_3;
hydrochloric acid, HCl

Fig. 61. Barium chloride solution and sodium carbonate solution before being mixed. *(Clara Aich and George Senty)*



What to do

PART I

1. Pour 10 ml of barium chloride solution into a 250-ml Erlenmeyer flask.
2. Pour 10 ml of sodium carbonate solution into a small test tube.
3. Carefully put the test tube into the flask. Stopper the flask. Be sure the stopper is tight. See Fig. 61.
4. Put the flask on a balance. Find the mass of the flask and its contents.

   The mass is ______________ grams.
5. Hold the stopper firmly in place. Tip the flask carefully so the liquids mix.

   Does a change take place? ______________

   Would you say that a chemical reaction occurred? ______________
6. Put the flask back on the balance. Find the mass.

   The mass is ______________ grams.

   Is mass conserved in this chemical reaction?

   ____

PART II

1. Get ready the apparatus and chemicals shown in Fig. 63. Use about 6 marble chips and 10 ml of 0.5*M* HCl.
2. Find the mass of the beaker and the test tubes and their contents.

   The mass is ______________ grams.
3. Add the acid to the marble chips. Put the beaker, with the two test tubes, back on the balance pan.

   What do you observe? ______________

   Find the mass when all reactions have stopped.

   The mass is ______________ grams.

Questions to answer

1. Was mass conserved in the reaction in PART I?

   ____

   What makes you think so? ______________

   ____
2. Do you think mass might be conserved in other chemical changes of this same type? ______________

   ____
3. Suppose you wanted to find if mass is conserved in the reaction in PART II. How would you do the experiment to find out? ______________

   ____

   ____

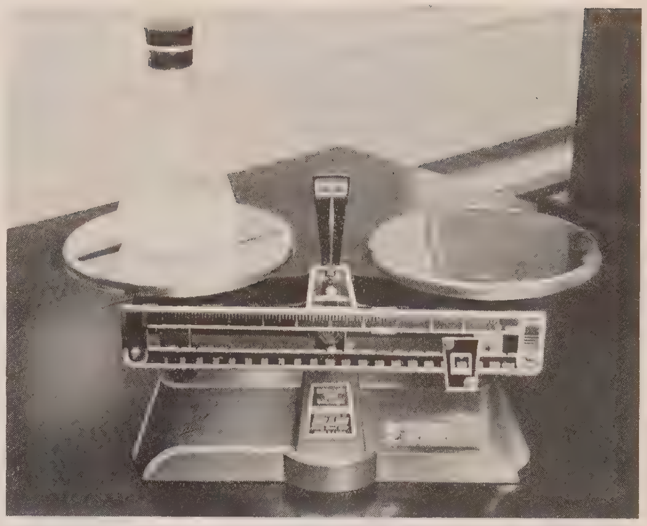


Fig. 62. After the chemical reaction between barium chloride solution and sodium carbonate solution. *(Clara Aich and George Senty)*

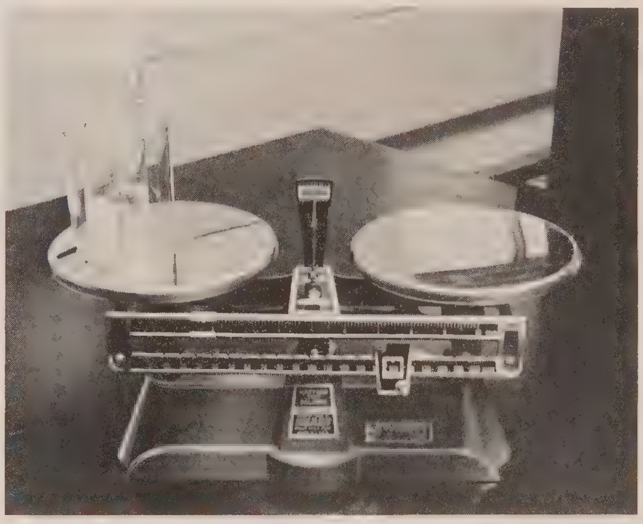
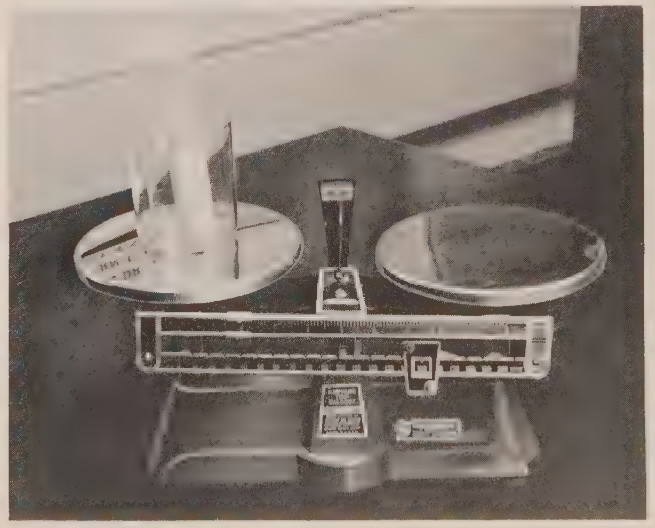
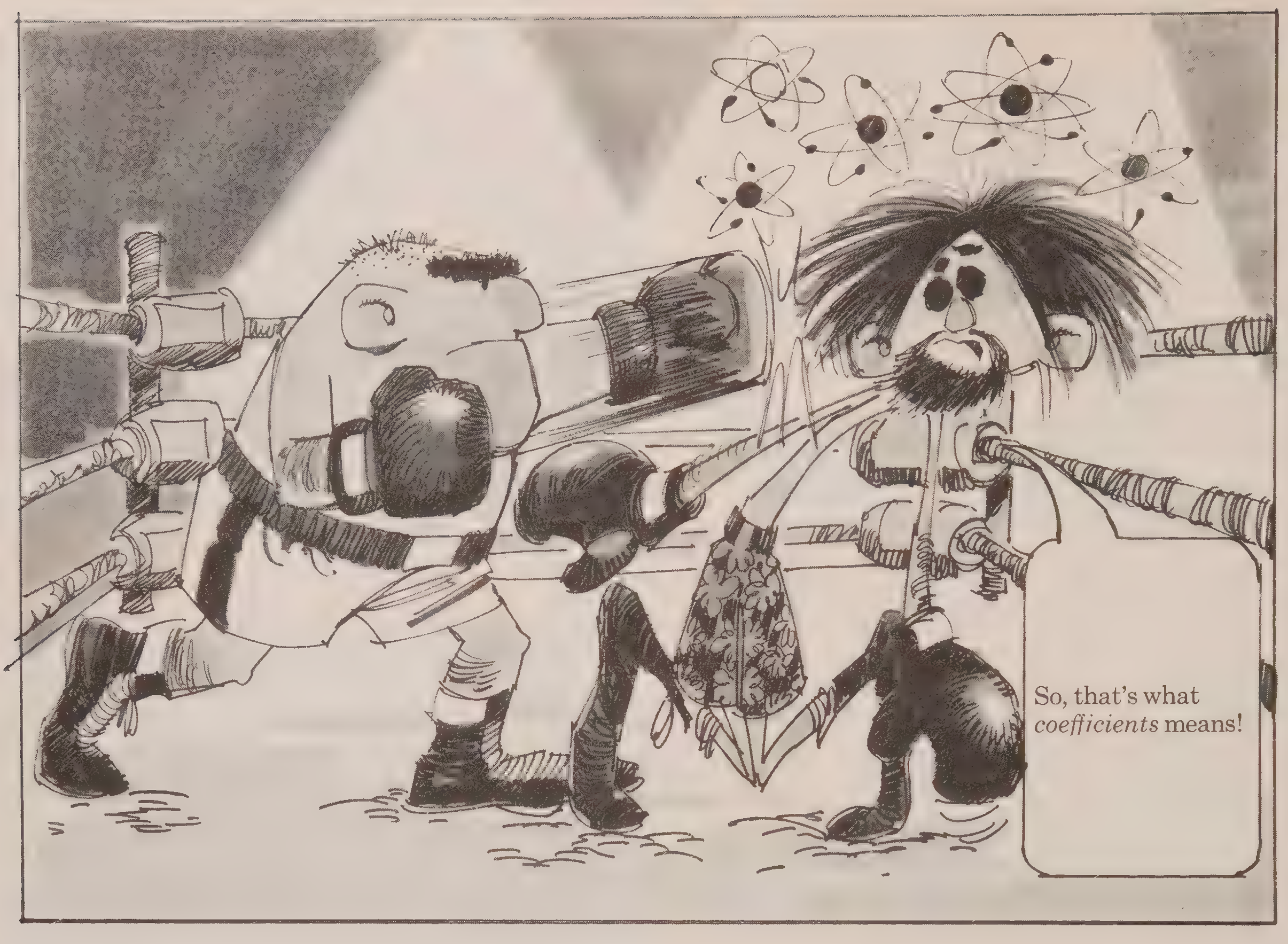


Fig. 63. Test tube containing marble chips and test tube containing hydrochloric acid solution. *(Clara Aich and George Senty)*

Fig. 64. Test tube containing marble chips and hydrochloric acid solution reacting, and empty test tube. *(Clara Aich and George Senty)*





So, that's what *coefficients* means!

Puzzles: Group and Regroup

A correct chemical equation gives us a lot of information. It tells us what has reacted and what has been formed. COEFFICIENTS help tell us the number of atoms taking part in the reaction. *COEFFICIENTS tell us the number of molecules in an equation.* Later, you will learn that balanced equations give us other information. First, you must learn to balance equations.

By solving the puzzles below, you will learn something about balancing equations. In the puzzles, you will use objects to stand for atoms. One or more objects might be grouped together. This grouping will be the smallest number of atoms of that substance that can exist by itself.

Remember:

1. An atom is the smallest part of an element that can take part in a chemical reaction.
2. A molecule is the smallest part of an element or compound that can exist by itself.

Experiment 13.2

Equipment and Supplies per Lab Station

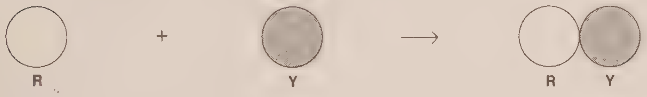
A puzzle kit containing:
(1) container of objects marked "R"
(1) container of objects marked "Y"
(2) "TAKE ONE AT A TIME" card
(2) "TAKE TWO AT A TIME" cards
(3) plus sign; (1) arrow; masking tape

PUZZLE RULES

1. Place the correct "Take" card before each container. The puzzle will tell you which card to use.
2. Take particles from the container to make the arrangement shown in each puzzle.
3. Fasten particles together by setting them down on the sticky side of masking tape.
4. Count the number of each kind of particle on each side of the arrow. If the numbers are not the same, you must add more so that you have the same number on both sides. This is making the sides "balance."

Puzzle 1

1. Place a "Take One at a Time" card before each container.
2. Find the number of each kind of particle needed to set up the following arrangement.



3. The number of R particles needed is ________.

   The number of Y particles needed is ________.
4. Below is the equation for this puzzle.

   $$R + Y \rightarrow RY$$

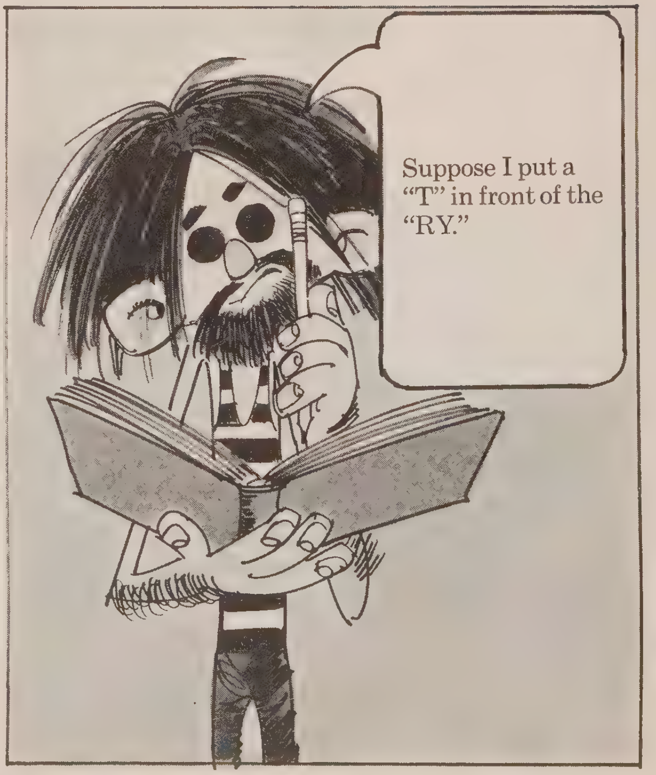
   (a) How many R particles are needed on the left side? ________
   (b) How many R particles are needed on the right side? ________
   (c) How many Y particles are needed on the left side? ________
   (d) How many Y particles are needed on the right side? ________
   (e) Write the correct number of particles needed in front of each letter.

   $$R + Y \rightarrow RY$$
5. Look at the following chemical equation. It describes the reaction between calcium and sulfur. When no coefficient is written before an element, that means only one atom is needed.

   $$Ca + S \rightarrow CaS$$

   Is it a balanced equation? ________

   How do you know? ________


Suppose I put a "T" in front of the "RY."

Is there a conservation of mass? ____________

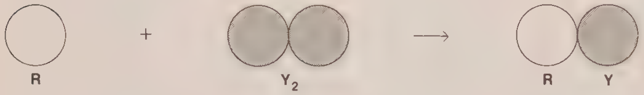
Why? ______________________________

____

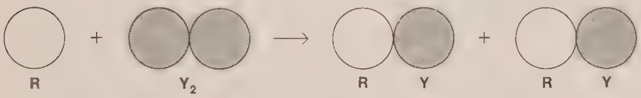
____

Puzzle 2

1. In front of the Y box of particles place the card which says "Take Two at a Time."
2. Follow the directions as you remove particles from the box. Set up the arrangement shown below.



3. If you followed directions, you will have one unused Y particle. But this unused particle must be part of the equation. Remember, you must have the same number of particles on one side as on the other. Put the extra Y particle on the right of the arrow. Since the product is RY add an R particle to it. You now have two RY particles on the right. The new arrangement looks like this.

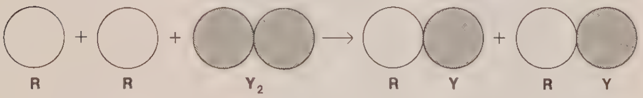


Why is this equation not balanced? __________
To balance the equation, add the necessary particle.

What particle did you add? ______________

Where did you add it? ______________

4. The arrangement you now have may look like this.



This can be written as an equation.

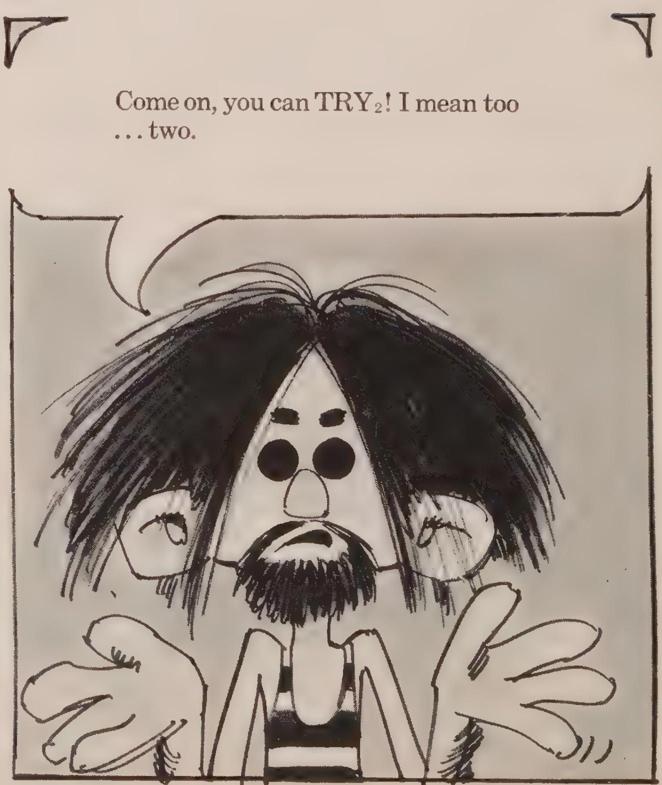
$$R + Y_2 \rightarrow RY$$

But the equation is not complete. To complete the equation, answer the following questions. Refer to the arrangement to help you.

(a) How many R particles are needed on the left side? ____________

(b) Write the number of R particles needed on the left next to the letter R.

(c) How many Y particles are needed on the left? ____________
Since you have that number already written as a subscript, you do not have to write anything.


Come on, you can TRY₂! I mean too . . . two.

(d) How many R particles should be on the right side? ________________

(e) How many Y particles should be on the right side? ________________

(f) Write the number of R particles needed on the right in front of the R. The R and Y are joined on the right side. The number you wrote applies also to the Y particle.

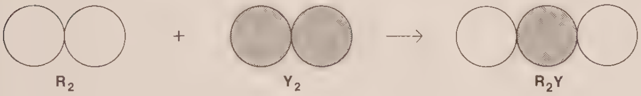
You should now have a balanced equation. Now try the same method for a chemical equation.

(g) Try to balance this chemical equation.

$$C + O_2 \rightarrow CO$$

Puzzle 3

1. Place the "Take Two at a Time" card before each container.
2. Follow the directions for removing particles. Set up the following arrangement.

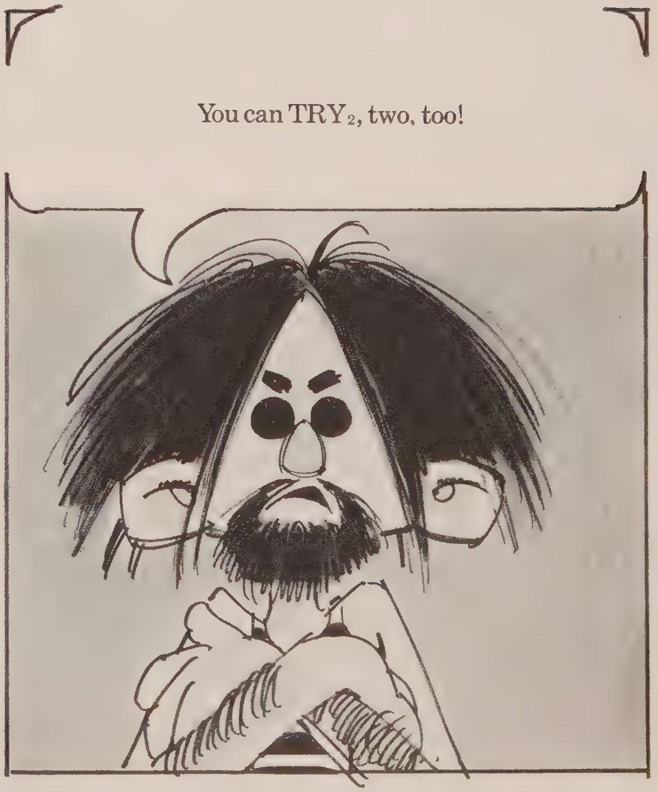


3. Again, you will have an unused particle. Use more particles until you can balance this arrangement. Be sure you follow directions: "Take Two at a Time."
4. When you have balanced the arrangement, write in the coefficients in the equation below. This shows how many groups you must have.

$$R_2 + Y_2 \rightarrow R_2Y$$

5. Now try to balance the following chemical equation.

$$H_2 + O_2 \rightarrow H_2O$$


You can TRY₂, two, too!

Questions to answer

1. Which of the following equations are balanced?

(a) $2R_2 + 3Y \rightarrow R_3Y_2$ ________________

(b) $3R_2 + 2Y \rightarrow 2R_3Y_2$ ________________

(c) $3R_2 + 4Y \rightarrow 2R_3Y_2$ ________________

2. Write the coefficient in front of each formula to balance the equation. Remember: a balanced equation will have the same number of atoms of each element on both sides of the arrow.

(a) $H_2 +$ $Cl_2 \rightarrow$ HCl

(b) $Ca +$ $O_2 \rightarrow$ CaO

(c) $Al +$ $S \rightarrow$ Al_2S_3

(d) $Al +$ $O_2 \rightarrow$ Al_2O_3

(e) $Zn +$ $Cl_2 \rightarrow$ $ZnCl_2$

Exercise 1: Balancing Equations

Balance each of the following equations. Put coefficients in front of formulas in order to obtain the same number of chemical units of each element on both sides of the arrow.

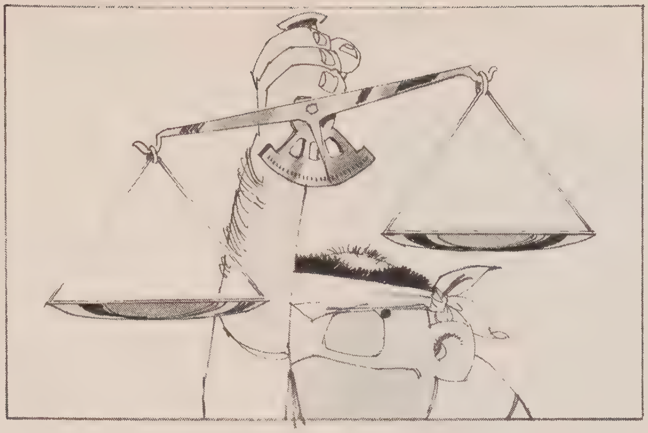
1. $Al + \quad S \rightarrow \quad Al_2S_3$
2. $Zn + \quad O_2 \rightarrow \quad ZnO$
3. $Mg + \quad Cl_2 \rightarrow \quad MgCl_2$
4. $Al + \quad O_2 \rightarrow \quad Al_2O_3$
5. $P + \quad O_2 \rightarrow \quad P_4O_{10}$
6. $Bi + \quad Cl_2 \rightarrow \quad BiCl_3$
7. $H_2 + \quad N_2 \rightarrow \quad NH_3$
8. $Cu + \quad O_2 \rightarrow \quad Cu_2O$
9. $Sn + \quad Cl_2 \rightarrow \quad SnCl_4$
10. $Na + \quad S \rightarrow \quad Na_2S$
11. $Ag + \quad I_2 \rightarrow \quad AgI$
12. $Pb + \quad O_2 \rightarrow \quad PbO$
13. $Al + \quad Br_2 \rightarrow \quad AlBr_3$
14. $Fe + \quad F_2 \rightarrow \quad FeF_3$
15. $Sn + \quad O_2 \rightarrow \quad SnO$
16. $Sb + \quad S \rightarrow \quad Sb_2S_5$
17. $Ca + \quad O_2 \rightarrow \quad CaO$
18. $Ba + \quad O_2 \rightarrow \quad BaO$
19. $Mg + \quad P \rightarrow \quad Mg_3P_2$
20. $K + \quad N_2 \rightarrow \quad K_3N$

Exercise 2: Balancing Equations

Balance each of the following equations. Put coefficients in front of formulas in order to obtain the same number of chemical units of each element on both sides of the arrow.

PART I

1. $Zn + \quad S \rightarrow \quad ZnS$
2. $Zn + \quad Cl_2 \rightarrow \quad ZnCl_2$
3. $Zn + \quad O_2 \rightarrow \quad ZnO$
4. $Mg + \quad S \rightarrow \quad MgS$
5. $Al + \quad S \rightarrow \quad Al_2S_3$
6. $Fe + \quad Cl_2 \rightarrow \quad FeCl_3$
7. $Al + \quad O_2 \rightarrow \quad Al_2O_3$
8. $Pb + \quad O_2 \rightarrow \quad PbO$
9. $Ag + \quad S \rightarrow \quad Ag_2S$
10. $K + \quad S \rightarrow \quad K_2S$



PART II

Balance each of the following equations:

1. $Zn + \quad HCl \rightarrow \quad ZnCl_2 + \quad H_2$
2. $NH_3 + \quad HCl \rightarrow \quad NH_4Cl$
3. $Al + \quad HCl \rightarrow \quad AlCl_3 + \quad H_2$
4. $Mg + \quad H_3PO_4 \rightarrow \quad Mg_3(PO_4)_2 + \quad H_2$
5. $Cu + \quad AgNO_3 \rightarrow \quad Cu(NO_3)_2 + \quad Ag$
6. $Ca + \quad Pb(NO_3)_2 \rightarrow \quad Pb + \quad Ca(NO_3)_2$
7. $Al + \quad Pb(NO_3)_2 \rightarrow \quad Pb + \quad Al(NO_3)_3$
8. $Zn + \quad Sn(NO_3)_4 \rightarrow \quad Zn(NO_3)_2 + \quad Sn$
9. $Cl_2 + \quad AlI_3 \rightarrow \quad AlCl_3 + \quad I_2$
10. $Br_2 + \quad CuI \rightarrow \quad CuBr + \quad I_2$
11. $NH_4OH + \quad FeCl_3 \rightarrow \quad NH_4Cl + \quad Fe(OH)_3$
12. $KBr + \quad Pb(NO_3)_2 \rightarrow \quad KNO_3 + \quad PbBr_2$
13. $AlCl_3 + \quad H_2SO_4 \rightarrow \quad Al_2(SO_4)_3 + \quad HCl$
14. $Al_2(SO_4)_3 + \quad BaCl_2 \rightarrow \quad BaSO_4 + \quad AlCl_3$
15. $Na_2CO_3 + \quad CaCl_2 \rightarrow \quad CaCO_3 + \quad NaCl$
16. $H_2SO_4 + \quad AlBr_3 \rightarrow \quad HBr + \quad Al_2(SO_4)_3$
17. $Na_3PO_4 + \quad BaCl_2 \rightarrow \quad Ba_3(PO_4)_2 + \quad NaCl$
18. $K_3PO_4 + \quad MgSO_4 \rightarrow \quad Mg_3(PO_4)_2 + \quad K_2SO_4$
19. $NaOH + \quad CuSO_4 \rightarrow \quad Na_2SO_4 + \quad Cu(OH)_2$
20. $Mg(OH)_2 + \quad H_2SO_4 \rightarrow \quad MgSO_4 + \quad H_2O$

Batter up!

Baseball players can be ranked on their ability to hit the ball. Metals can be ranked on their ability to replace another metal in a compound. In this experiment, you will study four metals and rank them according to their replacing ability.

Experiment 13.3

How do you like that! They've replaced my good old Louieville Slugger with one made of Al! And I don't mean American League, either! What are metals going to replace next?

Equipment and Supplies per Lab Station

graduated cylinder—25 ml or 50 ml; (3) test tubes
piece of sandpaper, about 5 cm × 5 cm
1 piece, about 2 cm × 3 cm, of each of these metals: copper, Cu; zinc, Zn; and lead, Pb
5 ml of each of the following 0.1*M* solutions: silver nitrate, $AgNO_3$; copper (II) nitrate, $Cu(NO_3)_2$; lead nitrate, $Pb(NO_3)_2$

What to do

1. Get three test tubes. Place 5 ml of silver nitrate solution in each.
2. Get three pieces of each of the following metals: copper, zinc, lead. If the metal samples are not bright, sandpaper them.
3. Put a metal sample into each test tube containing the silver nitrate solution.
4. Watch the test tubes for five minutes. Write your observations in the Data Chart.
5. Repeat steps 1–4, using fresh pieces of metal. But this time use copper nitrate solution instead of silver nitrate solution. Add your observations to the Data Chart.
6. Repeat steps 1–3, using fresh pieces of metal. This time use lead nitrate solution. Write your observations in the Data Chart.

DATA CHART

Metal	Silver Nitrate	Copper Nitrate	Lead Nitrate
Cu			
Zn			
Pb			

7. Study your Data Chart. Then try to answer these questions.
   (a) Which metal reacted with the greatest number of solutions? ____________
   Which metal reacted with the least number of solutions? ____________
   (b) List the metals in the following order.
   Most reactive ____________
   Second most reactive ____________
   Least reactive ____________
   (c) Where would you put silver in your list? ____
   Why would you put it there? ____________
   ____

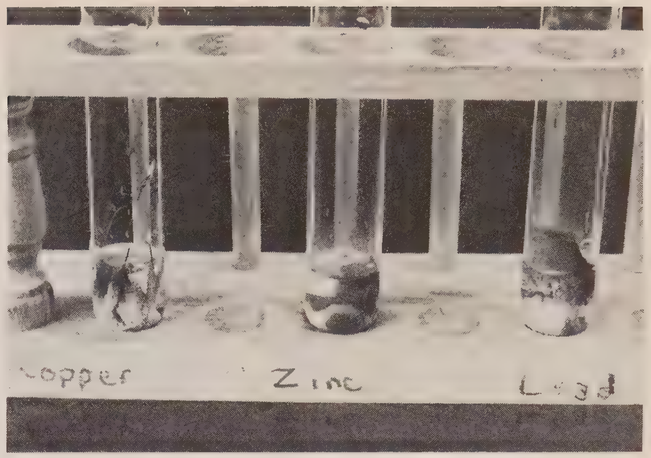


Fig. 65. *(Clara Aich and George Senty)*

Questions to answer

1. The reactions in this experiment are called single replacement reactions. Why do you think they are called this? ____________
   ____

2. Try to write a general rule describing single replacement reactions. ____________
   ____

3. Look at the ACTIVITY SERIES of metals found in reader p. 5/**19.**

   (a) List the following metals in order, with the most active at the top: Zn, Cu, Hg, Na.

   ____

   ____

   ____

   ____

   ____

   (b) Tell whether each of these reactions will or will not occur.

   $Zn + Mg(NO_3)_2 \rightarrow$

   $Cu + FeCl_3 \rightarrow$

   $Hg + AgNO_3 \rightarrow$

   $Na + Al_2(SO_4)_3 \rightarrow$

4. Write the balanced equation for each reaction that occurs in Question 3.

Holy cow, look at that. My team just lost 7,453,693 electrons!
TEAM A OXIDATION
We won! We won! My team just gained 7,453,693 electrons! I told you the REDUCTION team would always win!
TEAM B REDUCTION

Oxidation Reduction

In many chemical reactions, electrons are lost by a reactant. These lost electrons are gained by another reactant. The number of electrons gained always equals the number lost.

The losing of electrons by a reactant is called *oxidation.*

The gaining of electrons by a reactant is called *reduction.*

Reduction and oxidation happen together. You can't have one without the other. Many chemical reactions are oxidation-reduction reactions. These are sometimes called redox reactions.

Chemical equations are also used to describe oxidation-reduction reactions. The symbols used in oxidation-reduction equations are the same as those used for positive and negative ions. A symbol for a neutral atom is also used.

Here are some examples.

Na^0	An electrically neutral sodium atom
Na^+	A positively charged sodium ion
Cl^-	A negatively charged chlorine ion

In studying redox reactions, the idea of *oxidation number* is used. The oxidation number is a number that tells you how many electrons are lost or gained by an atom. Oxidation numbers are either positive (+) or minus (−). If a neutral atom loses electrons, it then becomes positively charged. It has a positive oxidation number. If a neutral atom gains electrons, it becomes negatively charged. It has a negative oxidation number.

An atom with an oxidation number of +2 means that it tends to lose two electrons. An atom with an oxidation number of −1 means that it tends to gain one electron.

For example, the oxygen atom almost always has an oxidation number of −2. Hydrogen has an oxidation number of +1.

In this experiment, you will study some oxidation-reduction reactions.

Experiment 13.3-A Extension

Equipment and Supplies per Lab Station

beaker—400 ml; evaporating dish
asbestos pad; burner; tongs
sandpaper, 3 cm × 3 cm; charcoal, powdered
copper metal foil, 3 cm × 2 cm
copper oxide, wire form; distilled water

What to do

PART I

1. Get a piece of copper about 3 cm by 2 cm. Sandpaper it until you can see the pure metal. How does the sandpapered metal differ from the original appearance? ______________________

   ____

   Try bending it a few times. Notice how it behaves. ______________________
2. Light a burner. Adjust it so that a blue cone is formed.
3. Hold the copper with a pair of tongs. Place it at the very top of the flame. In this position the oxygen of the air will touch the metal.
4. Heat the copper for about two minutes. Then place the copper on an asbestos pad so that it can cool. After it has cooled, look at it. How does it differ from the shiny, sandpapered copper? ______________________

   ____
5. Repeat steps 3 and 4 several times. Write down any changes that you observe.

   ____

   ____

6. Bend the copper back and forth. Does it act as the original copper did when you bent it? ______
If not, how does it differ? ______________

____

7. What evidence do you have that the original copper has reacted with oxygen in the air? ______

____

8. If copper reacted with oxygen, what substance was probably formed? ______________
9. Check this answer with a stockroom bottle of the substance. What compound did you form?

____

10. Balance the following equation.

$$Cu^0 + O_2^0 \rightarrow Cu^{+2}O^{-2}$$

(a) Did Cu^0 lose or gain electrons? __________
(b) Was Cu^0 oxidized or reduced? __________
(c) Did oxygen lose or gain electrons? ________
(d) Why is this reaction a direct union reaction? ______________

____

____

PART II

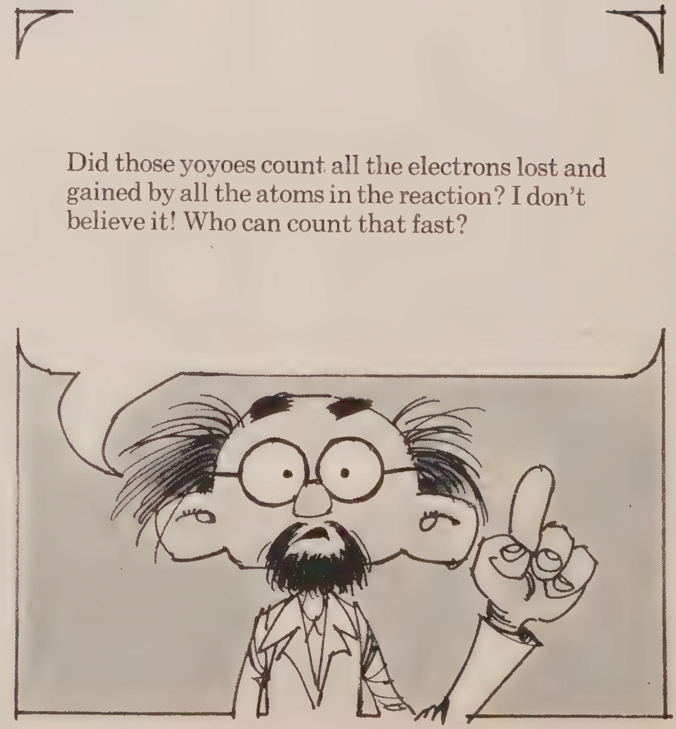
1. Mix a teaspoonful of powdered carbon with a teaspoonful of copper oxide.
2. Put the mixture in an evaporating dish. Place the dish on a ring stand.
3. Heat the mixture with a low flame for two minutes. Then increase the intensity of the heat. Heat it strongly for three minutes.
4. Cool the mixture.
5. Fill a beaker with about 250 ml of water. Put the cooled mixture into the water and stir gently.
6. Carefully pour off the black liquid. This removes the leftover carbon. Can you identify a new substance? ______________
If you cannot, add more water and stir. Pour off the water again. Repeat this until you can find a new substance.
7. Describe the new substance. ______________

____

What is this new substance? ______________
8. This equation describes the above reaction.

$$2\,CuO + C \rightarrow 2\,Cu + CO_2$$

(a) Place a zero at the top of any electrically neutral atoms of an element.
(b) Write the oxidation number for each substance in a compound.


Did those yoyoes count all the electrons lost and gained by all the atoms in the reaction? I don't believe it! Who can count that fast?

(1) What substance changed from zero to +4?

(2) Which changed from +2 to zero?

(3) Which substance gained electrons?

(4) Which substance lost electrons?

(5) What was oxidized?

(6) What was reduced?

(c) How do you know if this is an oxidation-reduction reaction?

(d) Why is this reaction a single replacement reaction?

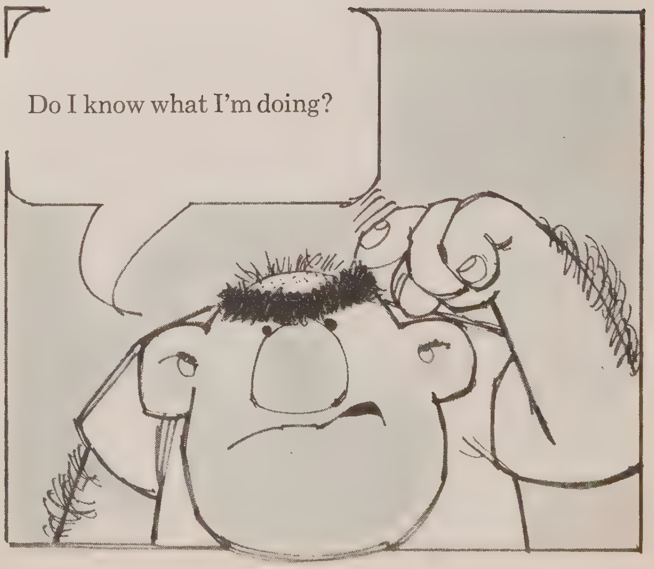
PART III

1. In Experiment 1.3 you placed aluminum foil into copper (II) chloride solution. The result was that aluminum gradually disappeared, the blue color of the copper chloride faded, and solid copper appeared in the liquid.
Write a balanced equation for that reaction.

2. (a) In the above reaction, did the copper ion gain or lose electrons?
How do you know?
(b) Did the chloride gain or lose electrons?
How do you know?
(c) Did the aluminum gain or lose electrons?
How do you know?
(d) Is this reaction an oxidation-reduction reaction?

Questions to answer

1. Can oxidation happen without reduction in a chemical reaction?
Explain your answer.
2. Describe how oxidation can happen without oxygen taking part.
3. Describe what happens in an oxidation-reduction reaction.


Do I know what I'm doing?

Chapter 14

Making Useful Compounds

Milk of magnesia is a common home remedy. It is also used in hospitals. Milk of magnesia is an antacid, or a laxative. Limewater is another commonly used substance. Limewater is often used in testing for carbon dioxide.

Both milk of magnesia and limewater are bases, and are used widely.
In this experiment, you will make these two common preparations.

Fig. 66. *(Clara Aich and George Senty)*

Experiment 14.1

Equipment and Supplies per Lab Station

beaker—150 ml; long medicine dropper; flask—125 ml
flask—250 ml; funnel; graduated cylinder—25 ml or 50 ml
balance; ring; ring stand; filter paper
litmus paper, red and blue
magnesium oxide, MgO, about .1 gram
calcium oxide, CaO, about 0.5 gram; distilled water
about 5 ml of 1*M* solutions of:
hydrochloric acid, HCl; sodium hydroxide, NaOH

What to do

PART I MAKING MILK OF MAGNESIA

1. Place one gram of magnesium oxide into a 125-ml flask.
2. Add 5 ml of distilled water to the magnesium oxide. Mix the water and magnesium oxide by swirling the flask.
3. Place 5 ml of 1*M* hydrochloric acid into a graduated cylinder.
4. Add the acid, drop by drop, to the magnesium oxide and water until all the magnesium oxide is dissolved. Use a long dropper to do this. How much acid was needed to dissolve the magnesium oxide? ______________ ml
5. Now add 1*M* sodium hydroxide to the flask containing the magnesium oxide. Add the same number of milliliters of sodium hydroxide as you used of the hydrochloric acid.
6. Put a stopper on the flask. Shake the solution for half a minute. Allow the solution to stand until the change has been completed.

   Describe the changes that you can see. ______
   ____
7. Carefully decant the clear liquid. You may throw away this liquid.
8. Add 10 ml of distilled water to the remaining solid. Shake it for about half a minute. Allow it to settle and again decant the liquid. Why do you think this step was done? ___________
   ____
   ____
   ____
   ____
9. Repeat the washing procedure of step 8 at least three more times.
10. Now add 10 ml of distilled water and shake well. Test with litmus paper. The solution should be slightly basic. If it is not, wash it until it is slightly basic.
11. When your solution is slightly basic, place it into a small bottle. Label it "Milk of Magnesia." Hand it into your teacher.

PART II MAKING LIMEWATER

1. Place 0.5 grams of calcium oxide into a 250-ml flask.
2. Add 100 ml of distilled water. Shake it well.
3. Let the mixture stand for about twenty minutes. Every five minutes give it a good shake.
4. Finally, swirl it to mix well. Pour it through a filter, a little at a time.
5. Collect the clear filtrate. Place it into a bottle and label it "Limewater."
6. Use this limewater to test for carbon dioxide. Do this by pouring about 10 ml of limewater into a test tube. Blow through a straw into the limewater. Blow so that bubbles are made continuously. Blow until a change happens.
   (a) What changes in the limewater do you observe? ______________________

   ____

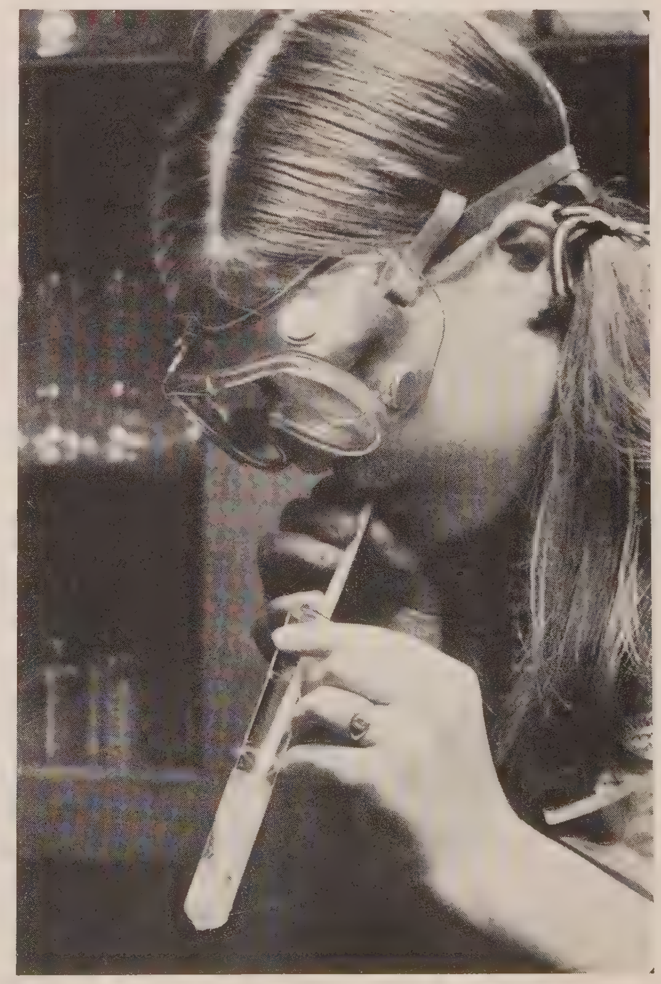
   (b) The chemical equation that describes this reaction is

   $$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$$

   Is the equation balanced? ______________
7. Hand the unused limewater to your teacher.

Limewater and a straw. To go.

Fig. 67. Student blowing carbon dioxide into limewater. *(Clara Aich and George Senty)*



EQUATIONS

The word equations for the chemical reactions in this experiment are:

Milk of Magnesia

magnesium oxide plus hydrochloric acid → magnesium chloride plus water

____

magnesium chloride plus sodium hydroxide → sodium chloride plus magnesium hydroxide

____

Limewater

calcium oxide plus water → calcium hydroxide

____

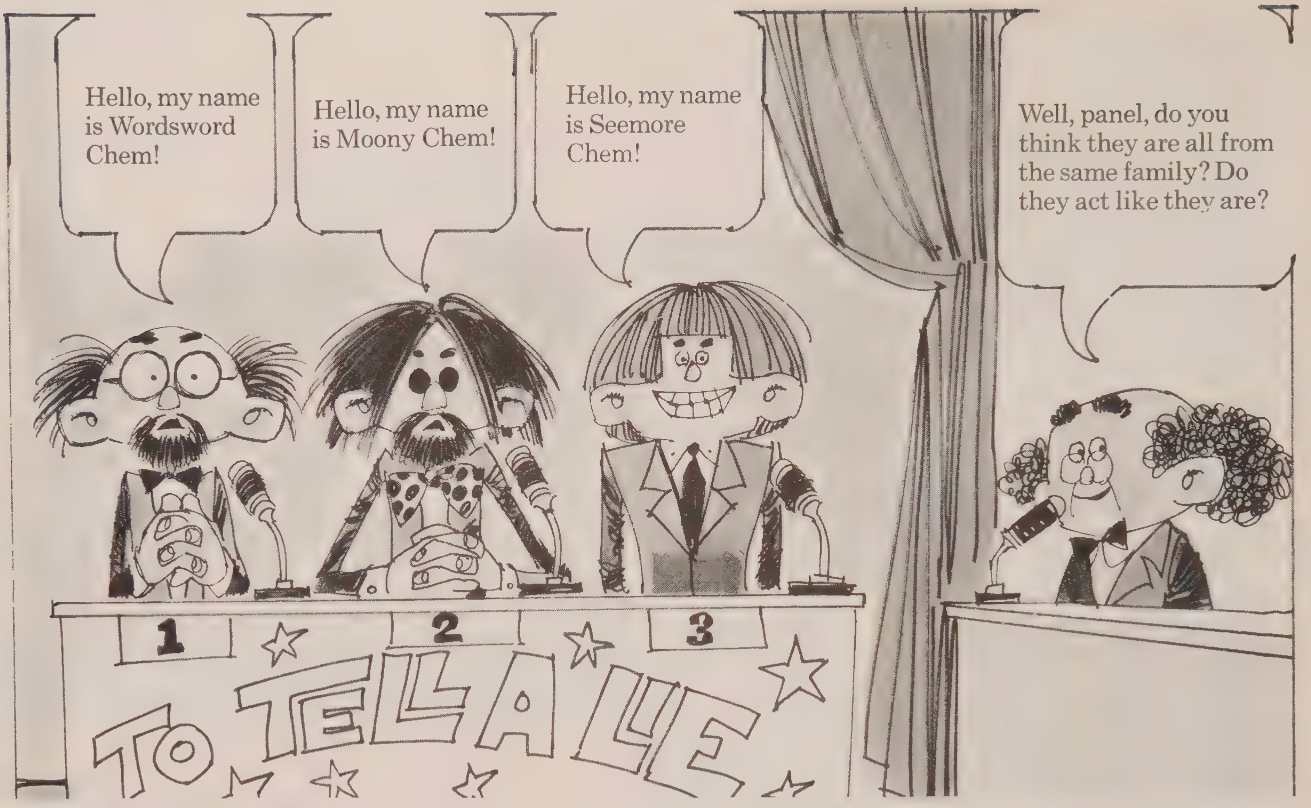
On the lines above:

1. Write the equations using chemical symbols.
2. Balance these equations.

CALCULATIONS

If you have already done Experiment 14.1, do the following calculations.

1. One mole of magnesium chloride yields ______________ moles of magnesium hydroxide.
2. One mole of calcium oxide yields ______________ moles of calcium hydroxide.


Hello, my name is Wordsword Chem!
Hello, my name is Moony Chem!
Hello, my name is Seemore Chem!
Well, panel, do you think they are all from the same family? Do they act like they are?
1
2
3
TO TELL A LIE

Family Resemblances

You have already studied some things about the Periodic Chart. You probably remember that the vertical columns are called families or groups. All the different elements in a family have something in common. They all have the same number of electrons in their valence energy level.

Strong family resemblances might be used to predict the chemical behavior of a family of elements. Suppose you know how one or two family members act in chemical reactions. Can you expect the other family members to behave in the same way?

In the experiment, you will try to answer that question.

Experiment 14.2

What to do

PART I GROUP II-A

1. Put 5 ml of 0.5 *M* magnesium nitrate, $Mg(NO_3)_2$, solution into a test tube. Into another test tube, place 5 ml of 0.5 *M* calcium nitrate, $Ca(NO_3)_2$, solutions.
2. Add 5 ml of 0.5 *M* sodium carbonate, Na_2CO_3, solution to each test tube. Shake each tube well.

   Describe what happens. ____________________

   ____

Equipment and Supplies per Lab Station

graduated cylinder—10 ml, 25 ml, or 50 ml
test tube rack; test tubes
5 ml of each of the following solutions:
0.5 *M* magnesium nitrate, $Mg(NO_3)_2$
0.5 *M* calcium nitrate, $Ca(NO_3)_2$
0.5 *M* strontium nitrate, $Sr(NO_3)_2$
0.5 *M* barium nitrate, $Ba(NO_3)_2$
0.2 *M* sodium chloride, NaCl
0.2 *M* sodium bromide, NaBr
0.2 *M* sodium iodide, NaI
0.5 *M* sodium carbonate, Na_2CO_3, 25 ml of the solution
0.2 *M* silver nitrate, $AgNO_3$, 20 ml of the solution

3. Strontium, barium, magnesium, and calcium are all members of the same family. Suppose you had solutions of strontium nitrate, $Sr(NO_3)_2$, and barium nitrate, $Ba(NO_3)_2$. Predict what would happen if you mixed each of those solutions with sodium carbonate solution. ____________________

   ____

   ____

4. Test your prediction by repeating the procedure in steps 1 and 2. Use 0.5 *M* strontium nitrate solution and 0.5 *M* barium nitrate solution.

   Describe the changes. ____________________

____

____

____

____

PART II GROUP VII-A

1. Place 5 ml of 0.2 *M* sodium chloride, NaCl, solution into a test tube.
2. Add 5 ml of 0.2 *M* silver nitrate, $AgNO_3$, solution to the sodium chloride solution.

   Describe what happens. ______________

   ____

3. Hold the test tube in a strong light. Sunlight is best if a window is nearby. Watch the contents of the tube for a few minutes. Describe any

   changes that you observe. ______________

   ____

4. Chlorine, bromine, and iodine are all members of the halogen family. Instead of using sodium chloride in the reaction, suppose you use sodium bromide, NaBr, or sodium iodide, NaI.

   Predict what you think would happen. ______

   ____

5. Check your prediction by repeating the procedure of steps 1, 2, and 3. First use a 0.2 *M* solution of sodium bromide. Then try a 0.2 *M* solution of sodium iodide.

   (a) Describe the changes. ______________

   ____

   (b) How did the color compare between the sodium bromide reaction and the sodium

   iodide reaction? ______________

   ____

   ____

EQUATIONS

1. Complete the word equations.
2. Beneath each word equation, write and balance the actual chemical equation.

PART I

(a) magnesium nitrate + sodium carbonate → ______________________

____

(b) calcium nitrate + sodium carbonate → ______________________

____

(c) strontium nitrate + sodium carbonate → ______________________

____

(d) barium nitrate + sodium carbonate → ______________________

____

PART II

(a) sodium chloride + silver nitrate → ______________________

____

(b) sodium bromide + silver nitrate → ______________________

____

(c) sodium iodide + silver nitrate → ______________________

____

Questions to answer

1. Suppose you were given a test tube containing a clear liquid. How could you find out if an

   element of Group II-A were in the liquid? ______________

   ____

2. How could you find out if your tap water has

   chlorine in it? ______________________

   ____

Moles and Equations

A recipe may say that one cup of pancake mix plus one cup of milk will make enough pancakes for two people. For one person then you would use one-half cup of each. For a crowd of forty-two people you would use twenty-one cups of pancake mix and twenty-one cups of milk.

The balanced chemical equation is the chemist's recipe for making chemical products. It tells him how much of the reactants he will need to prepare the desired amount of the product. It also tells how much of a product he can expect from a starting quantity of one of the reactants. These predictions have been proved reliable in countless numbers of experiments.

Information given by the balanced chemical equation

The balanced equation for reaction between iron and sulfur is:

$$Fe + S \rightarrow FeS$$

First of all, this means that one atom of iron combines with one atom of sulfur to form one molecule of iron (II) sulfide. If one atom combines with one atom, then 6.02×10^{23} atoms of iron will combine with 6.02×10^{23} atoms of sulfur to form 6.02×10^{23} molecules of iron (II) sulfide. One mole of a substance contains 6.02×10^{23} atoms or molecules of the substance. Therefore, one mole of iron combines with one mole of sulfur to form one mole of iron (II) sulfide.

The balanced chemical equation gives the chemist all of the following information:

$$\begin{array}{ccccc} Fe & + & S & \rightarrow & FeS \end{array}$$

1 atom of iron + 1 atom of sulfur → 1 molecule of iron(II) sulfide

1 mole of iron + 1 mole of sulfur → 1 mole of iron(II) sulfide

56 g/mole of Fe + 32 g/mole of S → 88 g/mole of FeS

88 g of reactants → 88 g of product

Using the information in the balanced equation

Example: A chemist has 112 g of iron.

1. How many grams of sulfur will combine completely with 112 g of iron?

(a) 112 g of iron $= \frac{112\text{ g}}{56\text{ g/mole}}$ moles of iron $= 2$ moles of iron

(b) 1 mole of iron combines with 1 mole of sulfur

$$Fe + S \rightarrow FeS$$

(c) If twice as many moles (2 moles) of iron are used, then twice as many moles of sulfur will be needed.

$$2\ Fe + 2\ S \rightarrow 2\ FeS$$

(d) 2 moles of sulfur $= 2 \times 32$ g/mole $= 64$ g of sulfur needed

2. How many grams of iron(II) sulfide can be made from 112 g of iron?

(a) 112 g of iron $= \frac{112\text{ g}}{56\text{ g/mole}}$ moles of iron $= 2$ moles of iron

(b) 1 mole of iron combines with 1 mole of sulfur to produce 1 mole of iron(II) sulfide.

$$Fe + S \rightarrow FeS$$

(c) If twice as many moles (2 moles) of iron are used, then twice as many moles (2 moles) of iron(II) sulfide will be produced.

$$2\ Fe + 2S \rightarrow 2\ FeS$$

(d) The mass of one mole of FeS is found as follows:
56 g/mole of Fe + 32 g/mole of S = 88 g/mole of FeS

(e) The mass of two moles of FeS then would be:
2 moles of FeS $= 2 \times 88$ g/mole $= 176$ g
Then 176 g of FeS can be made from 112 g of iron.

PROBLEMS

1. The balanced equation for the decomposition of aluminum oxide, Al_2O_3, is

$$2\ Al_2O_3 \rightarrow 4\ Al + 3\ O_2$$

How many grams of aluminum oxide, Al_2O_3, would be needed to produce a 216 gram roll of aluminum foil?

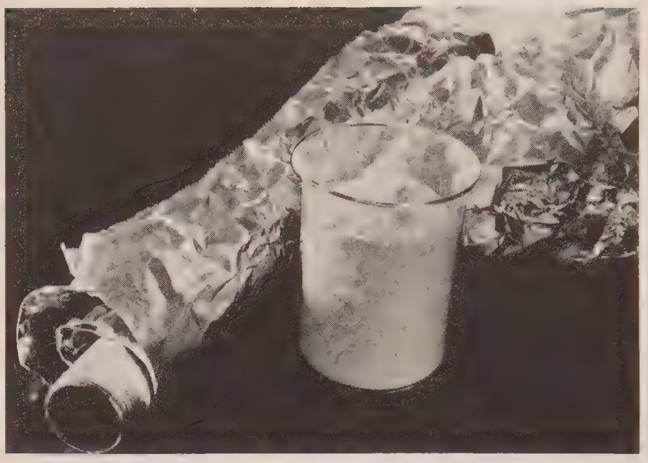


Fig. 68. Two hundred and sixteen grams of aluminum and the alumina that produces it. *(Clara Aich and George Senty)*

2. Magnesium is used in making space ships because of its low density. The salt, magnesium chloride, is the second most common salt in seawater. Magnesium chloride extracted from seawater is used in the production of magnesium.

How much magnesium chloride, $MgCl_2$, would be required to make a nose cone using 4,400 kilograms of magnesium? (Hint: divide each formula weight in the equation by 1000 to change it from grams to kilograms.) Equation:

$$MgCl_2 \rightarrow Mg + Cl_2$$

3. One mole of a gas occupies 24.5 liters at sea level pressure (760 mm of mercury) and room temperature (25°C).

In a new antipollution device, how many liters of oxygen are needed to change 24.5 liters of deadly carbon monoxide, CO, gas from an automobile exhaust into harmless carbon dioxide, CO_2, gas?

First balance the equation, then work out the answer.

$$CO + O_2 \rightarrow CO_2$$

4. The bleaching power of hydrogen peroxide, H_2O_2, is due to oxygen released according to the following equation:

$$2\ H_2O_2 \rightarrow O_2 + 2\ H_2O$$

If 3.2 grams of oxygen will "frost" a particular lock of hair, how many grams of hydrogen peroxide will be needed to produce the 3.2 grams of oxygen?

5. How many grams of silver can be electroplated onto a copper teaspoon from a solution containing 1700 grams of silver nitrate, $AgNO_3$?

First balance the equation, then work out the problem.

$$AgNO_3 + Cu \rightarrow Ag + Cu(NO_3)_2$$

How Much Chalk Can You Make?

Calcium carbonate is an abundant chemical, and it is found in different forms. One type of calcium carbonate is chalk. (Not the blackboard chalk that you know.) Chalk is made for use as a gentle scouring powder. For that reason, it is used in toothpaste.

Chalk is easy to make. You merely combine calcium chloride and sodium carbonate. In fact, you can learn how to make just the amount you want, and no more. Suppose you were given known amounts of calcium chloride and sodium carbonate. How much chalk could you make?

Experiment 14.3

Equipment and Supplies per Lab Station

(2) beaker-250 ml; beaker—150 ml; funnel
graduated cylinder—25 ml or 50 ml
stirring rod; balance; ring; ring stand
filter paper; distilled water
sodium carbonate, Na_2CO_3,-2.5 grams
20 ml of 1 *M* solution of calcium chloride, $CaCl_2$

What to do

1. Label two, clean, dry, 250-ml beakers, respectively 1, 2.
2. Fold a piece of filter paper to make a filter. Place it in Beaker 1. Find the mass of the beaker and filter paper.
   Mass of Beaker 1 + filter paper: __________ g
3. Set up a funnel for filtering. Now place the weighed filter paper into the funnel. Dampen it slightly with distilled water.
4. Place 20 ml of 1.0 *M* calcium chloride, $CaCl_2$, solution in Beaker 1. Set it aside where it will be safe.
5. Place 2.5 g of sodium carbonate, Na_2CO_3, into Beaker 2.
6. Add 20 ml of distilled water to Beaker 2. Swirl the beaker until all of the sodium carbonate has dissolved.
7. Pour the sodium carbonate solution into the 20 ml of calcium chloride solution. Swirl gently to mix. Describe the changes that happen. ______________________________

   ____
8. Pour the mixture down a stirring rod into the filter. Add a little at a time. Each time fill the filter not more than half full. Do not spill any.

9. Watch the filter until all the liquid has drained through. Now wash the solid in the filter by adding about 5 ml of distilled water. Let the wash water drain through. Repeat the washing twice more. Save the filtrate and wash water. You may need it.
10. Gently loosen the filter paper from the funnel. Place the filter paper and its contents into the 250-ml beaker labelled 1. Set the beaker and its contents aside to dry overnight.
11. In the next laboratory period, find the mass of Beaker 1 and its contents.
    Beaker 1 + filter paper + product ________ g

CALCULATIONS

1. Mass of Product

   Beaker 1 + filter paper + product ________ g

   Beaker 1 + filter paper (from step 1) ________ g

   Mass of product ________ g

2. Moles of Calcium chloride, $CaCl_2$
   You used 20 ml of a 1 *M* solution of calcium chloride. The 1 *M* solution contained one mole of calcium chloride in one liter, 1000 ml, of solution.
   Therefore, 20 ml of 1 *M* calcium chloride contains a fraction of one mole.

   $$1 \text{ mole} \times \frac{20 \text{ ml}}{1000 \text{ ml/liter}} = \_\_\_\_\_\_\_\_ \text{ mole}$$

   of $CaCl_2$.

3. The Chemical Equation
   This equation describes the reaction in this experiment.

   $$CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + 2\,NaCl$$

   (a) Both calcium carbonate and sodium chloride can be solids. Look up the solubility of each of these products. Which one is probably the solid precipitate in this case? ________
   (b) Which one remains in solution as ions?

   ____

4. Moles in the Balanced Equation
   (a) According to the balanced equation

   ________ mole of $CaCl_2$ + ________ mole of Na_2CO_3 →

   ________ mole of $CaCO_3$ + ________ mole of NaCl

Arrright, who took my moles?

(b) In your experiment, ______________ mole(s) of $CaCl_2$ reacted with Na_2CO_3. How many moles of the product would you expect to form? ______________ mole(s) of product.

5. Moles of Product Produced
   (a) Calculate the formula mass of your product. Remember, this is the number of grams/mole.

   Formula mass of product = ________ g/mole.
   (b) Calculate the moles of the product.

   Moles of product = grams of product

   $$\times \frac{1\ \text{mole}}{\text{grams/mole}} = \_\_\_\_\_\_\_\_\_\_\_\_\ \text{mole(s)}$$

6. Checking Theory Against Experiment
   (a) How many moles of the product did you expect? This was calculated from the equation, problem 4 (b) above.

   ______________ mole(s)
   (b) How many mole(s) of the product were actually made? This is your answer to problem 5 (b) above. ______________ moles
   (c) Did you get more moles or fewer moles than the equation predicted? ______________
   (d) What was the difference in the number of moles predicted ___ and the number made?

7. Grams of Product in Your Experiment
   (a) Using the equation, find the number of grams of the product that is predicted. Moles of product (your answer to 4 (b) above)

   ______________ moles

   Grams of product = ______________ mole(s)

   × ________ g/mole = ________ g predicted
   (b) How close were you to the predicted results?

   grams predicted by the equation ________g

   grams actually produced ________g

   difference ________g
   (c) Finding the percent difference

   $$\frac{\text{grams actually made in the experiment}}{\text{grams predicted by the equation}} \times 100$$

   = ______________% difference.

____

____

____

Question about the experiment

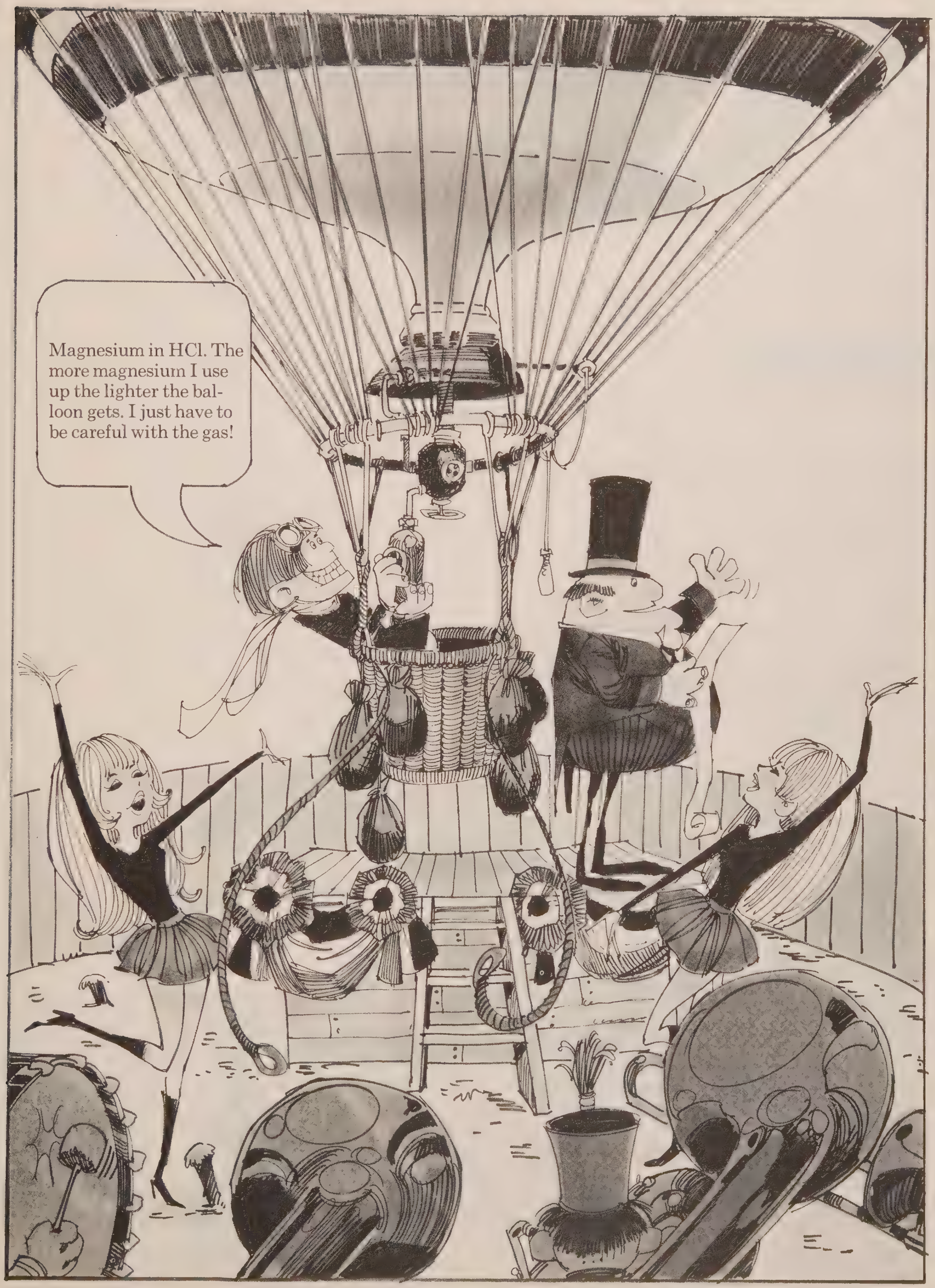
If you have over 10% difference between your calculated result and the result your obtained in laboratory, can you explain it?

____

____

____

____


Magnesium in HCl. The more magnesium I use up the lighter the balloon gets. I just have to be careful with the gas!

Trading Magnesium for Hydrogen

Hydrogen can be made by the reaction of magnesium with hydrochloric acid. You would expect that the more magnesium you use, the more hydrogen you get as long as there is an excess of hydrochloric acid. Suppose you doubled the amount of magnesium used. Would you double the volume of hydrogen produced? In this experiment, you will try to answer that question.

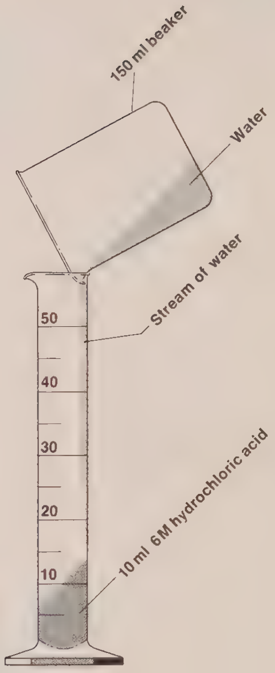


Fig. 69.

Experiment 14.4

Equipment and Supplies per Lab Station

evaporating dish—250 ml
graduated cylinder—50 ml
clamp; ring stand; cotton
(2) pieces of magnesium ribbon, Mg, ($2^{1}/_{4}$ cm long) and ($4^{1}/_{2}$ cm long)
10 ml of 6 *M* hydrochloric acid solution, HCl

What to do

1. Get two pieces of magnesium ribbon, one 2-1/4 cm long and one 4-1/2 cm long.
2. Wrap each piece of magnesium into a loose coil. Place the 2-1/4 cm magnesium into a 250-ml evaporating dish. Pour about 150 ml of water into the evaporating dish.
3. Read steps 4, 5, and 6. Practice them with water before you add acid.
4. Carefully pour 10 ml of 6 *M* hydrochloric acid into a 50-ml graduated cylinder. Pour the water very slowly into the graduated cylinder. Fill it completely. See Fig. 69.
5. Place a ring stand and clamp next to the evaporating dish.
6. Cover the top of the cylinder with a wad of wet cotton. Hold the cotton with your fingers and turn the cylinder upside down. Place the mouth of the cylinder directly over the magnesium ribbon. Be sure there are no air bubbles in the cylinder. Remove the cotton. Support the cylinder with a clamp. See Fig. 70.
7. Describe the changes that are happening. ____

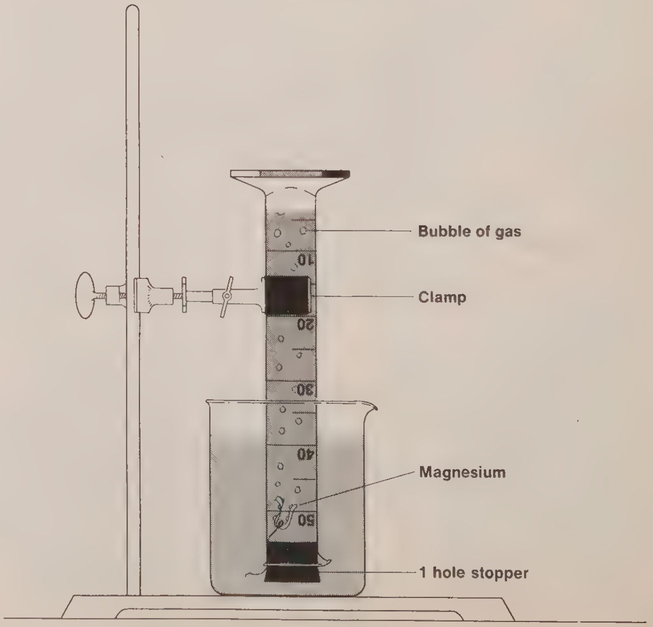
____

____

____

____

Fig. 70.



8. After the reaction is finished, read the volume of the gas produced. Be sure to keep your eye level with the meniscus.
   The volume of hydrogen gas produced is

   ______________ ml
9. Repeat the experiment. This time use the 4-1/2 cm strip of magnesium.
   What is the volume of hydrogen produced?

   ______________ ml
10. Complete the following table.

Amount of Magnesium Used	Volume of Hydrogen Produced
2-1/4	______________ ml
4-1/2	______________ ml

Questions to answer

1. In this experiment, you doubled the amount of magnesium used. Did the volume of hydrogen double or almost double? ______________
2. Suppose you would use a 6-3/4 cm magnesium strip. How much hydrogen do you think you would get? ______________
3. Write the balanced equation for the reaction between magnesium and hydrochloric acid.

   ____

   ____
4. Look at the balanced equation. What clue is there that might explain the volume of hydrogen? ______________

   ____

A Closer Look at the Magnesium-Hydrogen Switch

In Experiment 14.4, you learned that magnesium reacts with hydrochloric acid to make hydrogen gas. You discovered that if you used more magnesium, you made more hydrogen.

In this experiment, you will study the mole relationship between two measurements: the mass of magnesium used and the volume of hydrogen produced.

Experiment 14.4-A Extension

Equipment and Supplies per Lab Station

barometer for class; beaker—400 ml
graduated cylinder—50 ml; thermometer
buret clamp; ring stand; cotton
magnesium ribbon, Mg, about 4.5 cm
10 ml of 6 *M* hydrochloric acid solution, HCl

What to do

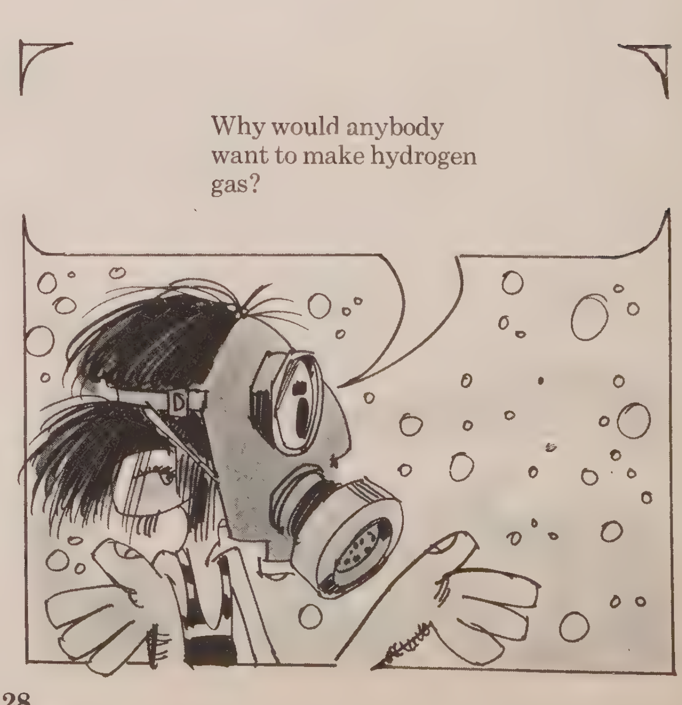
1. Measure and cut a piece of magnesium ribbon 4.5 cm long.
2. From your teacher, find the mass of 100 centimeters of magnesium ribbon. This is ______________ grams.
3. Now calculate the mass of your piece of magnesium ribbon.

   $\frac{4.5 \text{ cm}}{100 \text{ cm}} \times \text{grams} = \text{mass of your sample}$

   Mass of your magnesium is ______________ grams.
4. Other information is also needed in this type of experiment.
   (a) What is the temperature of the room? ______________°C
   (b) Ask your teacher for the barometric pressure. ______________mm of mercury
   (c) Use the Vapor Pressure Table to find the vapor pressure of water at room temperature. ______________mm of mercury

VAPOR PRESSURE OF WATER AT VARIOUS TEMPERATURES

Temperature °C	Pressure (mm of Mercury)	Temperature °C	Pressure (mm of Mercury)
15	12.8	23	21.0
16	13.6	24	22.4
17	14.5	25	23.8
18	15.5	26	25.2
19	16.5	27	26.7
20	17.5	28	28.3
21	18.6	29	30.0
22	19.8	30	31.8


Why would anybody want to make hydrogen gas?

A Note About Air Pressure

The pressure of the air is important when studying gases. The air pressure, or barometric pressure, changes with altitude and weather conditions. The higher you go, the less the pressure. The hotter a parcel of gas, the more pressure it exerts. Air pressures can be measured in millimeters of mercury. At *sea level*, the pressure of the air will push up a column of mercury *760 mm high.*

5. Fill in the following Data Chart.

DATA CHART

Mass of magnesium used ______________________ g

Room temperature ______________________ °C

Barometric pressure ______________________ mm

Vapor Pressure of Water at Room Temperature ________ mm

Barometric Pressure at Sea Level ______________________ mm

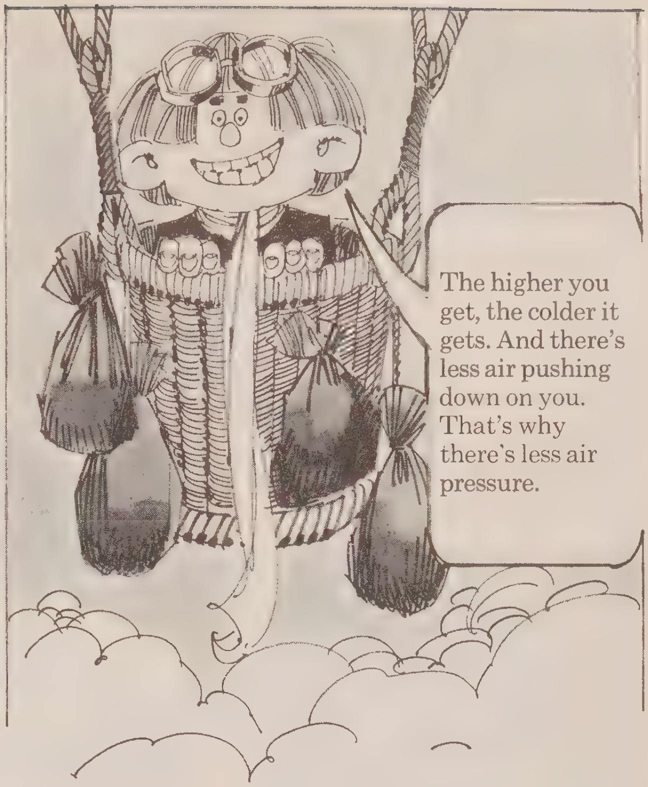
6. Now wrap your magnesium strip into a coil. Place it into a 400-ml beaker. Pour about 250 ml of water into the beaker.
7. Read steps 8, 9, and 10. Practice them with water before you try them with acid.
8. Carefully pour 10 ml of 6 *M* hydrochloric acid into a 50-ml graduated cylinder. Pour water very slowly into the cylinder. Fill it completely.
9. Place a ring stand and clamp next to the beaker.
10. Cover the top of the cylinder with a flat wad of wet cotton. Hold the cotton in place with your fingers. Turn the cylinder upside down. Place the mouth of the cylinder directly over the magnesium ribbon. Be sure there are no air bubbles in the cylinder. Remove the cotton. See Fig. 70 in Experiment 14.4. Support the cylinder with the clamp.
11. Watch the reaction until all of the magnesium has disappeared.
12. Now carefully loosen the clamp. Raise or lower the cylinder until the water level inside the cylinder is the same as it is in the 400-ml beaker. Add water to the beaker if necessary.
13. Read the volume of the hydrogen gas. Make sure your eye is level with the meniscus in the cylinder. The volume of the gas is

    ______________ ml.

Making the Calculations

1. *The Pressure of Dry Hydrogen*

   You probably wondered why you had to move the graduated cylinders to make the water levels the same (step 12). In doing that, you made the pressure of the gas in the cylinder the same as the outside barometric pressure.


The higher you get, the colder it gets. And there's less air pushing down on you. That's why there's less air pressure.

(a) The pressure of the atmosphere is

______________ mm.

The pressure of the trapped gas is

______________ mm.

But not all of the gas inside the cylinder is hydrogen. Some of it is water vapor. Water vapor also exerts a pressure. To find the pressure of the water vapor, check the Data Table of Vapor Pressures.

(b) The vapor pressure is ______________ mm.

Therefore, to find the pressure made only by the hydrogen, you must subtract the vapor pressure. Pressure of trapped gas − vapor pressure equals hydrogen gas pressure

(c) ______________ mm of pressure of trapped gas equals

______________ mm vapor pressure minus

______________ mm of pressure of dry hydrogen gas

2. *Moles of Hydrogen Made in Your Experiment*
A mole of gas takes up quite a bit of space. Suppose the pressure of the atmosphere is 760 mm of mercury and the temperature is 25°C. Then a mole of gas would take up 24.5 liters of space.

You measured the volume of your hydrogen gas in milliliters. A milliliter is only a thousandths part of a liter. It takes 1000 milliliters to make a liter. For example, if your hydrogen volume was 32 milliliters in volume, what would the volume be in liters? To find out, you divide the 32 ml by 1000. This gives

$$32 \text{ ml} \times \frac{1 \text{ liter}}{1000 \text{ ml}} = \frac{32}{1000} \text{ liter or, } 0.032 \text{ liter}$$

(a) What is the volume of your collected hydrogen gas in liters? ________________

$$\text{vol. in ml} \times \frac{1 \text{ liter}}{1000 \text{ ml}} = \_\_\_\_\_\_\_\_\_\_\_\_$$

liter(s) of dry H_2 gas produced

Now let's find out how many moles of gas you collected. To do this, divide the volume collected by 24.5 liters. Suppose you collected 0.032 liters. This would be

$$0.032 \text{ liters} \times \frac{1 \text{ mole}}{24.5 \text{ liters}} = 0.014 \text{ moles}$$

(b) How many moles of hydrogen gas did you collect?

$$\text{vol. collected} \times \frac{1 \text{ mole}}{24.5 \text{ liters}} =$$

________________ moles of gas collected.

3. *Moles of Magnesium Used*
(a) What mass of magnesium did you use?

________________ g.

(b) A mole of magnesium is 24 grams of magnesium. To find the number of moles of magnesium used, you must divide by 24.

________________ grams of magnesium

$\frac{1 \text{ mole}}{24 \text{ g}}$ equals ________________ moles of magnesium

4. *Moles of Hydrogen Made by One Mole of Magnesium*
The balanced equation for this reaction is

$$Mg + 2\,HCl \rightarrow H_2(\text{gas}) + MgCl_2$$

(a) Look at the equation. Is one mole of magnesium traded for one mole of hydrogen?

____

How do you know? ________________________

(b) Write the moles of magnesium used and moles of hydrogen produced in your experiment.

________________ $Mg + 2\,HCl$

________________ $H_2(\text{gas}) + MgCl_2$

(c) In your experiment, was it an even trade, one mole of hydrogen for one mole of magnesium? ________________________

(d) Try to describe the relationship between moles of magnesium and moles of hydrogen.

____

____

Questions for Discussion

1. Suppose you had used aluminum instead of magnesium in this experiment.
(a) Write the balanced equation for the reaction. Hint: Remember that aluminum becomes the Al^{+++} ion.

$\_\_\_\_Al + \_\_\_\_HCl \rightarrow \_\_\_\_H_2 + \_\_\_\_\_\_\_\_$

(b) Would one mole of aluminum have been traded for one mole of hydrogen? ____________.

(c) How do you know? ________________________

2. In this experiment, you changed the grams of magnesium used to moles. You also changed the milliliters of hydrogen produced to moles also. Why do you think this was done? ____________

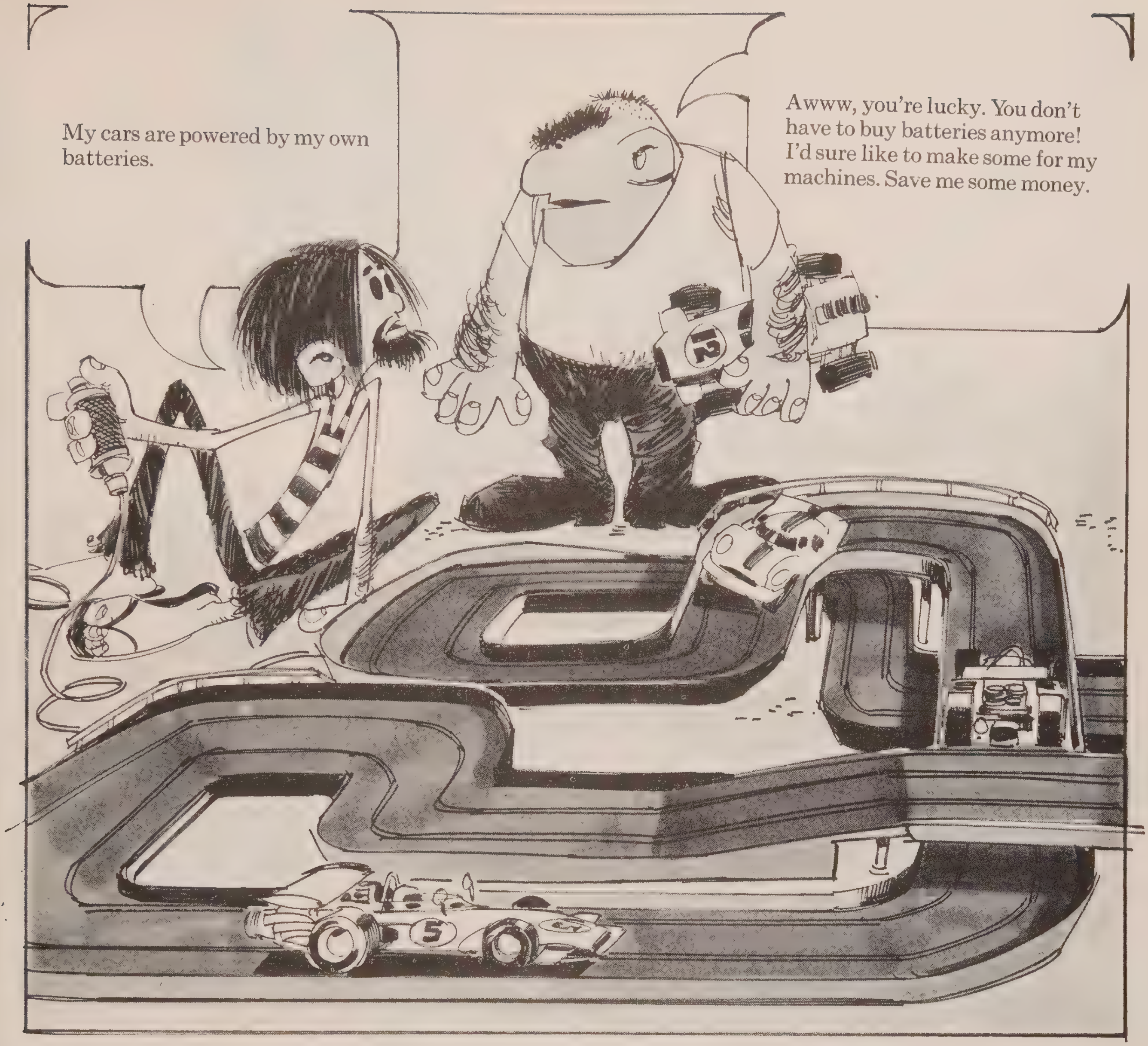
____

3. Suppose you wanted to fill a balloon with a large amount of hydrogen gas. To get enough gas, you would have to figure out the amounts of metal and acid to use. Would you do this the same way you did in this experiment? ____________

Describe how you might do it. ________________________

____

____


My cars are powered by my own batteries.
Awww, you're lucky. You don't have to buy batteries anymore! I'd sure like to make some for my machines. Save me some money.
12
5

Making Batteries

Experiment 14.5

Batteries work because of chemical reactions. There are different sizes and strengths of batteries. In this experiment, you will make two simple batteries. By testing them, you will decide which one is the stronger.

Equipment and Supplies per Lab Station

beaker—150 ml; voltmeter; rubber band
(2) insulated copper wires, each with two alligator clips
(2) pieces of aluminum foil, Al, heavy duty, (12 in x 18 in)
piece of copper foil, Cu, (½ in x 3 in)
piece of lead foil, Pb, (½ in x 3 in)
80 ml of 0.1 *M* hydrochloric acid, HCl
80 ml of 0.05 *M* sulfuric acid, H_2SO_4

What to do

PART I

1. Cut a piece of heavy duty aluminum foil, 12 inches by 18 inches. Fold it in half so that you have a piece 12 inches by 9 inches.

2. Mold this double thickness over a 150-ml beaker as shown in Fig. 71. Trim the top edges. You have made an aluminum can.
3. Now fit the aluminum can inside the 150-ml beaker. Fold the top edge down over the top of the beaker.
4. Fill the aluminum can about half full of 0.1 *M* hydrochloric acid.
5. Stretch a rubber band over the top of the beaker. Run it under the beaker. See Fig. 72.
6. Hook a strip of copper onto the rubber band so that it hangs down into the hydrochloric acid. *The copper must not touch the aluminum.*
7. Get two insulated copper wires with alligator clips. Attach one end of one wire to the aluminum can by means of the alligator clip.
8. Attach the other wire to the copper strip by means of the alligator clip. See Fig. 73.
9. Attach the wire from the aluminum to the *negative* terminal of a voltmeter. Attach the wire from the copper to the *positive* terminal of the voltmeter.
10. What is the reading on the voltmeter?

_______________ volts

A voltmeter can measure the strength of a battery. In PART II you will make a different battery.

PART II

11. Repeat steps 1, 2, and 3. Fill this aluminum can about half full with 0.05 *M* sulfuric acid.
12. Suspend a strip of lead in the acid. *Be sure the lead does not touch the aluminum.* Use the copper wires to connect the lead strip to the positive terminal of the voltmeter.
13. Attach the aluminum can to the negative terminal of the voltmeter.
14. What is the reading on the voltmeter?

_______________ volts

Questions to answer

1. Did your batteries have the same strength? ____

If not, how much difference was there? ________

____

Which battery was stronger? __________________

2. Try to explain the operation of your two batteries. Hint: Refer to Experiment 13.3. _____

____

3. Try to explain why one battery was stronger.

____

____

____

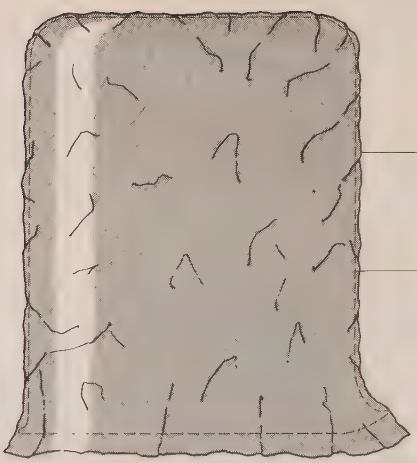
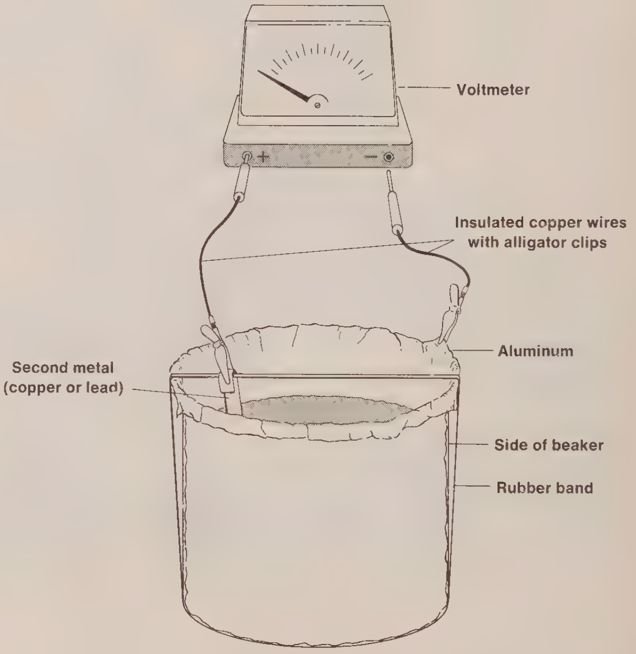


Fig. 71.

Fig. 72.


Voltmeter
+
−
Insulated copper wires with alligator clips
Aluminum
Second metal (copper or lead)
Side of beaker
Rubber band

Extra for experts

1. Look up the construction of a car battery.

Describe how it works. ______________________

____

____

____

____

____

____

2. What kind of tests are made on a car battery at a service station? ______________________

____

Laboratory Experiments in

ACTION CHEMISTRY

UNIT VI

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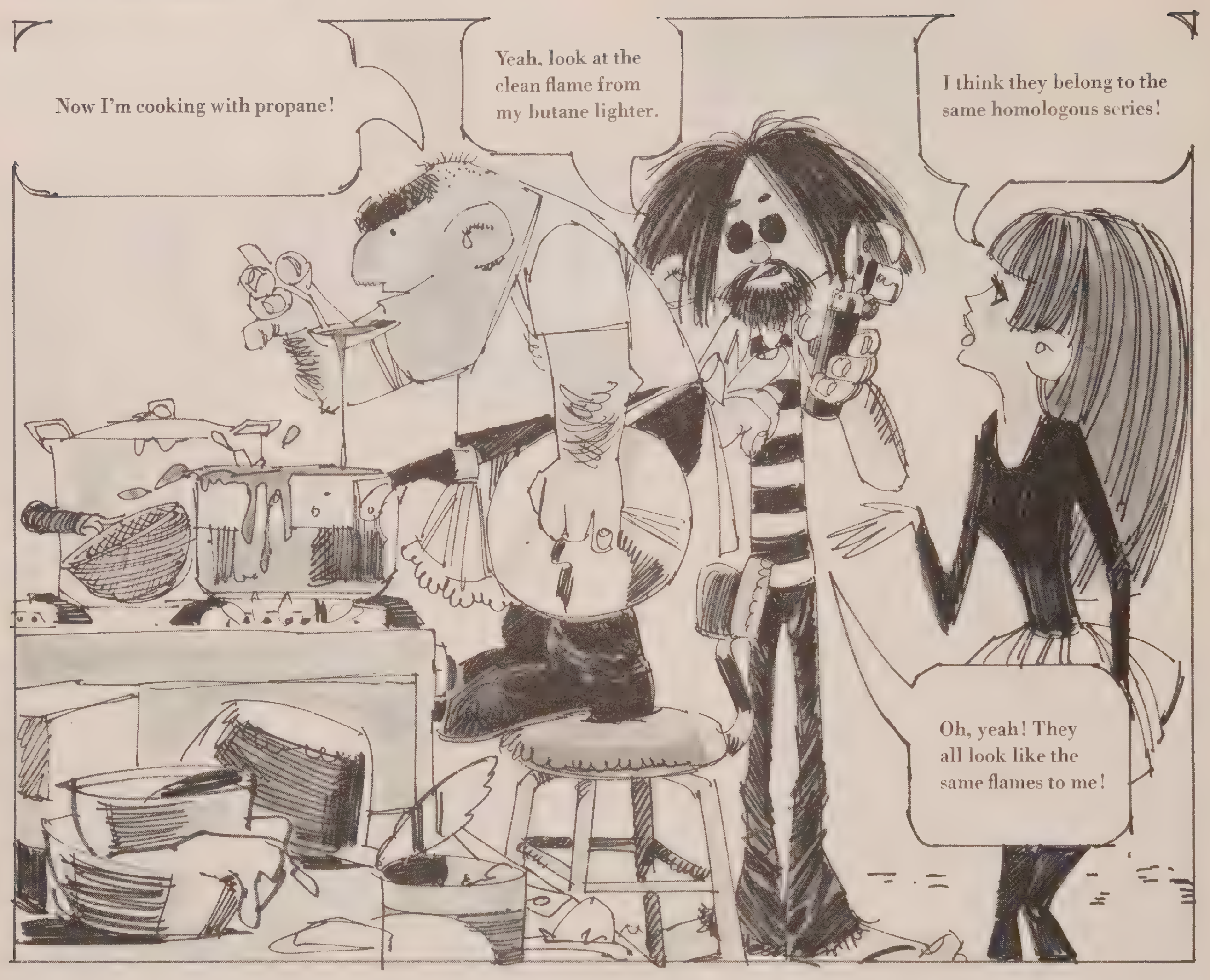

Now I'm cooking with propane!
Yeah, look at the clean flame from my butane lighter.
I think they belong to the same homologous series!
Oh, yeah! They all look like the same flames to me!

THE HOMOLOGS

A carbon atom may form covalent bonds with other carbon atoms. It is possible for compounds with either short or long chains of carbon atoms to be formed. Compounds with carbon chains may become an HOMOLOGOUS SERIES.

An HOMOLOGOUS SERIES is a series of similar compounds which conform to a general formula. (Remember, "homo" means "alike".)

In this experiment you will build models for molecules with carbon chains. You will find out whether compounds called alkanes, others called alkenes, and those called alcohols form homologous series. Also, you will discover that alcohols contain a FUNCTIONAL GROUP.

A FUNCTIONAL GROUP is that part of the molecule that does the work or enters into characteristic reactions.

Equipment and Supplies per Lab Station
ball and stick molecular model kit

What to do

1. Obtain a ball and stick model kit. The instruction sheet in the box lists the color code for the balls. Let one ball represent an atom. A stick is used to show one covalent bond. Springs are used to represent two or more bonds between the same two atoms.

2. Make a model of methane, CH_4. Record the structural formula in Data Chart 1.

3. Find successive members of the series in which the HYDROCARBON, methane, is the first member. *(A HYDROCARBON is a compound composed of hydrogen and carbon only.)*
   (a) Replace one hydrogen atom in methane or one end hydrogen atom on a carbon chain with another carbon atom.
   (b) Add as many more hydrogen atoms as the new carbon atom can take.
   (c) Record the formulas in Data Chart 1.
   (d) Repeat Steps (a), (b), and (c) until all the Data Chart 1 information has been found.

DATA CHART 1

ALKANE SERIES			
Member	Molecular Formula	Structural Formula	Name
First	CH_4		methane
Second			ethane
Third			propane
Fourth			butane
Fifth			pentane

4. Study Data Chart 1 and answer the following questions:
   (a) For every additional carbon atom, how many hydrogen atoms are needed? ______________
   (b) The consecutive members of the homologous series differ by $-CH_2$. Does ethane belong to an homologous series? ______________

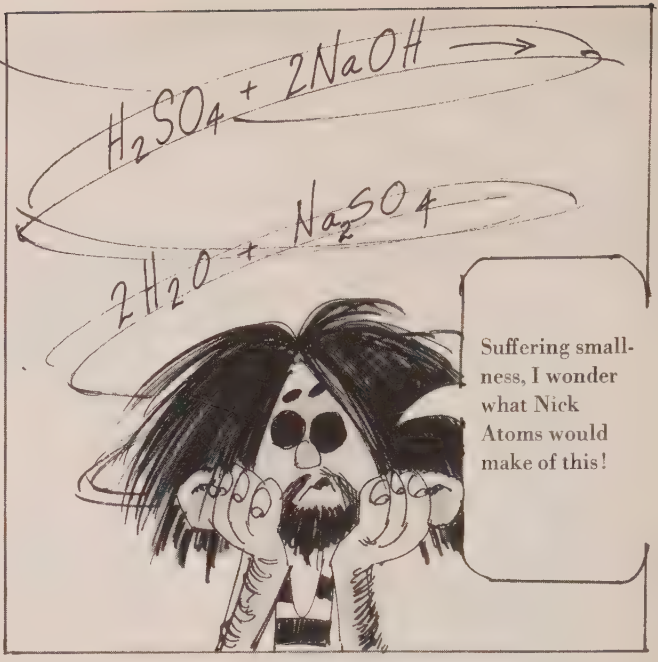
(c) The members of an homologous series have a mathematical relationship between the number of carbon atoms and the number of hydrogen atoms.

Example 1: There may be twice as many hydrogen atoms as carbon atoms. This relationship would be written as C_nH_{2n}. The expression is called a GENERAL FORMULA. The subscript n stands for the number of carbon atoms.

Example 2: If the hydrogen atom count is two more than twice the number of carbon atoms, the general formula would be written as C_nH_{2n+2}.
What does the general formula C_nH_{2n-2} mean?

____

____


$H_2SO_4 + 2NaOH \rightarrow$
$2H_2O + Na_2SO_4$
Suffering smallness, I wonder what Nick Atoms would make of this!

(d) Does the Alkane Series (also called the Methane Series) fit any one of the above formulas? ______________ If so, which general formula? ______________________

(e) Can you predict the formula for an alkane having an eight carbon atom chain? __________ If so, what is its formula? ____________________

5. Select two carbon atoms and connect them by two springs. Complete the model by adding as many hydrogen atoms as possible. The model represents a molecule of ethene, also called ethylene, the first member of the Alkene Series. Record the structural formula in Data Chart 2.

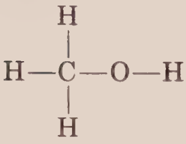
6. Find out if the Alkene Series is an homologous series. Repeat step 3. Keep the pair of double-bonded carbon atoms at one end of the molecular model. The other carbon bonds are single bonds. Continue to add carbon atoms until all of the information for Data Chart 2 has been found.

DATA CHART 2

ALKENE SERIES			
Member	Molecular Formula	Structural Formula	Name
First	C_2H_4		ethene (ethylene)
Second			propene (propylene)
Third			butene
Fourth			pentene

7. Study the Data Chart 2 and answer the questions:
   (a) Are the members of the Alkene Series members of a homologous series? ______________ How do you know? ______________________________
   ____
   (b) What is the general formula of the Alkene Series? ______________________________

8. Make a model for a methanol molecule. Methanol is an alcohol whose structural formula is



$H-\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-O-H$. The —O—H is known as a functional group; it is found in all alcohols.

9. Repeat step 3 to lengthen the carbon chain. Keep the —O—H group in all of your models. Write the structural formulas for the models you make.

10. Can the alcohols you made models of be a homologous series? ______________________________
   Explain. ______________________________
   ____

Questions to answer

1. How is the presence of one double bond indicated in the name of a homologous series? ______________

2. Could you have a homologous series with the general formula C_nH_{2n-2}? ______________ If so, how would you bond the carbon atoms together? ______________
   ____

3. Could you have alcohols with more than one —O—H group? ______________________________

ISOMERS

Two builders can use the same amount of materials but construct two houses with different structures. Chemists can take the same amounts of the same elements and bond them together in different arrangements. The different arrangements produce compounds which have completely different properties.

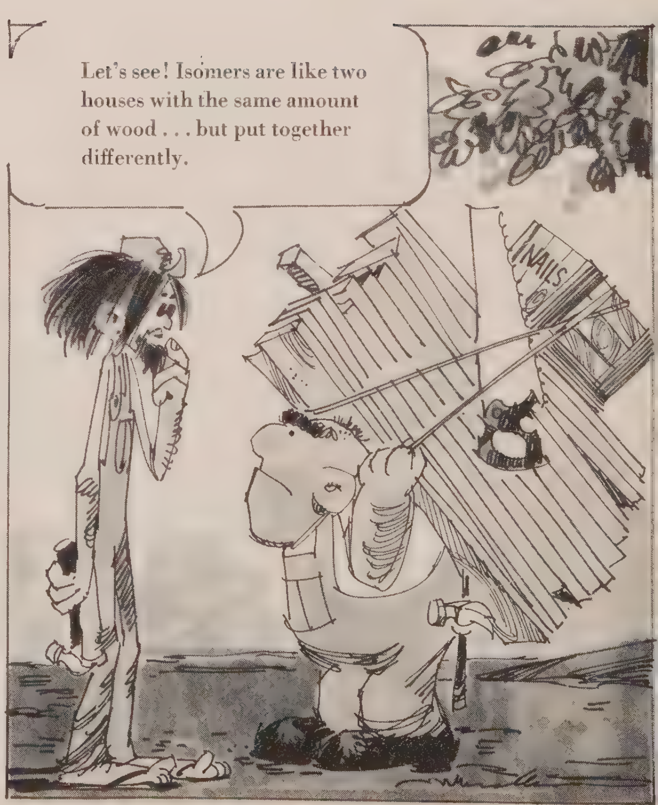
Molecules which have the same number of atoms of each element but different arrangements of atoms are called ISOMERS.

Experiment 15.2

Equipment and Supplies per Lab Station
ball and stick molecular model kit

What to do

1. Obtain a molecular model kit.
2. Make a model of the butane molecule, C_4H_{10}.
3. Write the structural formula for the model made in step 2.
4. Set the model aside for comparison later with another model (step 5).
5. Select 4 carbon atoms and 10 hydrogen atoms. Put them together but not in the same way as you did in step 2.
6. Write the structural formula for the model you made in step 5.


Let's see! Isomers are like two houses with the same amount of wood . . . but put together differently.
NAILS

7. Are the two models isomers? ____________________

Why or why not? ____________________

____

8. Select 2 springs, 10 sticks, 4 carbon atoms, and 10 hydrogen atoms. See how many different arrangements you can make using all of these materials. Write in the space at the right the structural formula for each isomer you make.

9. Make two different arrangements using two carbon atoms, six hydrogen atoms, and one oxygen atom. Use as many sticks and/or springs as necessary. Write in the space at the right the structural formula for each isomer.

10. Make a model of pentane, C_5H_{12}. Remove one hydrogen atom from an end carbon atom. Replace it with the —O—H functional group of an alcohol.

11. Draw in the space at the right the structural formula of the alcohol molecule (pentanol).

12. Make two isomers of the pentanol molecule. Write in the space at the right the structural formula for each isomer.

Questions to answer

1. Given the formulas $CH_3CH_2OCH_2CH_3$ and $C_2H_5COC_2H_5$, how can you tell whether these substances are isomers? ____________________

____

2. Are CH_3COOCH_3 and CH_3COOH isomers? ____________________

How do you know this? ____________________

____

3. State the ways in which isomers may differ. (Look at your structural formulas.) ____________________

____

____

ALCOHOLS VS BASES

Organic compounds containing the hydroxl group, —OH, are classified as alcohols. Inorganic substances known as hydroxides contain the hydroxide ion, OH^-. Are alcohols like bases (hydroxides)? Compare them and find out.

Experiment 15.3

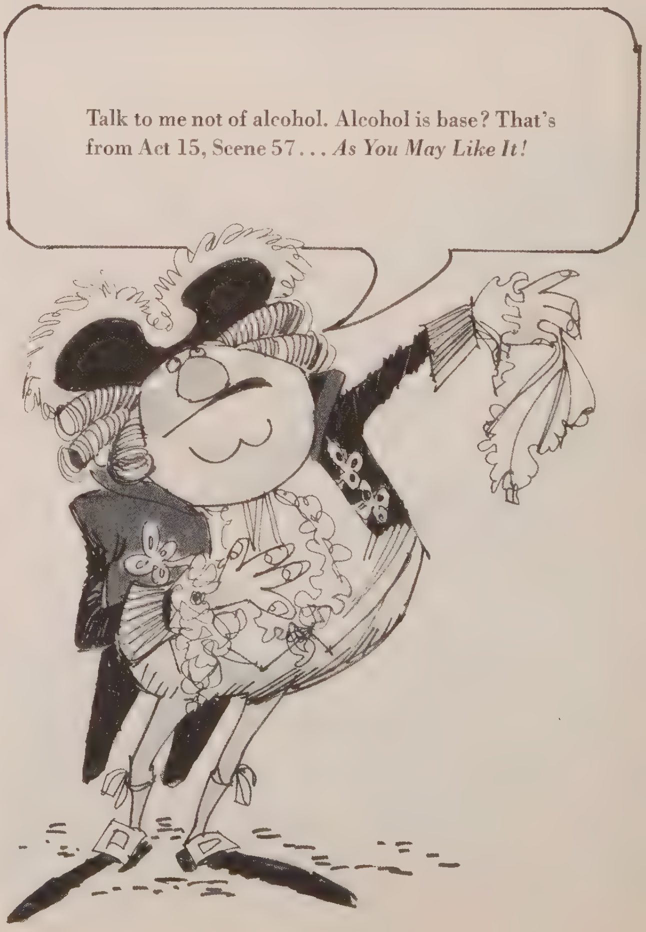
Equipment and Supplies per Lab Station

graduated cylinder—50 ml; stirring rod, glass; test tubes
conductivity apparatus with small beakers to fit
forceps; spoon; litmus paper
sodium hydroxide, NaOH, solid flakes or pellets
calcium hydroxide, $Ca(OH)_2$, powdered
barium hydroxide, $Ba(OH)_2$, powdered
ethanol; glycerol; isopropyl alcohol (propanol-2)

What to do

1. Inspect samples of the following:
   Alcohols: ethanol, propanol-2, and glycerol
   Bases: calcium hydroxide, barium hydroxide, and sodium hydroxide
   List the phase of each substance at room temperature in the Data Chart on page *6*/**9**.
2. Test the solubility of each alcohol by pouring 10 ml of the alcohol into 10 ml of water in a test tube. Stir. Write your observations in the Data Chart.
3. Test the mixtures with red litmus paper. Save the materials for use later.
4. Set up a conductivity apparatus to test the electrical conductivity of the liquid.
5. Test each alcohol and its solution with the apparatus. Add your observations to the Data Chart.
6. Be sure to use a spoon or forceps when working with the bases. Do not inhale their dust or touch them with your hands. Immediately wash with quantities of water any area where any base or basic solution gets spilled.
7. Repeat step 2 but use one pellet or 1/4 spoonful of powdered base. Write your data in the the chart.
8. Repeat steps 3 and 4. Test only the solutions of the bases. Add your observations to the Data Chart.
9. Study the Data Chart. Are alcohols and bases the same kind of compounds? ________________ Why or why not? ____

____


Talk to me not of alcohol. Alcohol is base? That's from Act 15, Scene 57 . . . *As You May Like It!*

DATA CHART

Substance	Phase	Solubility in H_2O	Effect on Red litmus	Electrical conductor	
				liquid	solution

Questions to answer

1. Do you think that the bonding is the same in bases and alcohols? ______________ What makes you think this? ______________________________

2. Look up the structural formulas of ethanol, propanol-2, and glycerol. Draw them here.

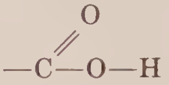
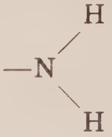
3. What differences do you see among the above structures? ______________________________

____

____

Chemical Chain Gangs

Experiment 15.4

Amino acids are organic compounds. Their molecules contain two functional groups. One group is the *carboxyl group,*  $-\overset{O}{\overset{/\!/}{C}}-O-H$ which is found in all organic acids. The second functional group is the *amino group,*  $-N\genfrac{}{}{0pt}{}{\diagup H}{\diagdown H}$ which may also be found in amines.

Amino acids are the simple molecules from which giant protein molecules can be formed. *The process by which the amino acid molecules join to form the proteins is called POLYMERIZATION.* How does polymerization occur? The answer can be found in this experiment.

Equipment and Supplies per Lab Station
ball and stick molecular model kit

What to do

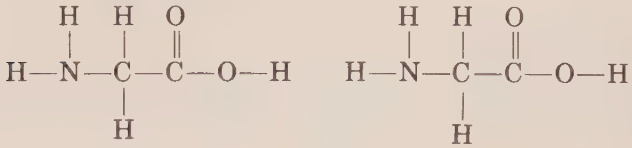
1. Obtain a ball and stick molecular model kit.
2. Make an amino functional group. Use three sticks, two hydrogen atoms, and one nitrogen atom. Make it just as illustrated above.
3. Set the amino functional group aside for later use.
4. Make a carboxyl functional group. Set it aside for later use.
5. Make a model of methane, CH_4.
6. Remove one hydrogen atom from the methane molecule. Replace it with the carboxyl group. What kind of compound does the new model represent?

   _______________ Write the structural formula for your model.

7. Remove another hydrogen atom from the methane carbon atom. Replace it with the amino group. Look at the functional groups in the Exercise. What kind of compound does your new model represent? _______________ Write the structural formula for your model.

8. Make another model like the one made in step 7.

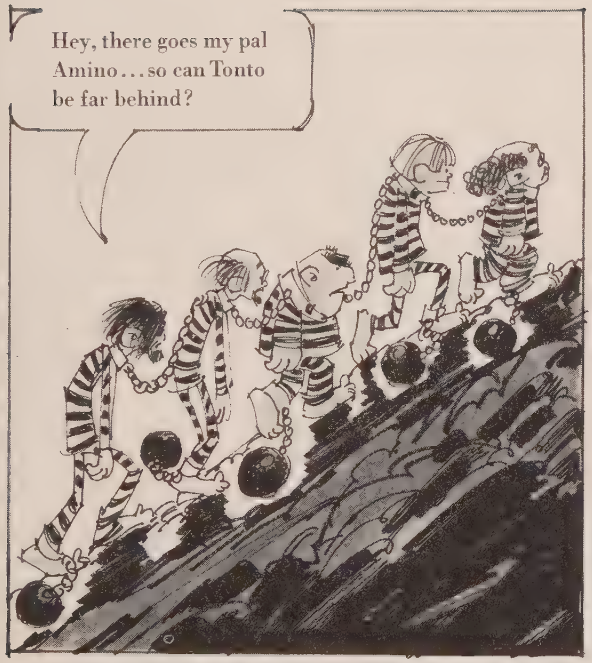
9. Arrange the models in steps 7 and 8 so that the carboxyl group of one model faces the amino group of the other model.



$$H-\overset{H}{\overset{|}{N}}-\underset{\underset{H}{|}}{\overset{H}{\overset{|}{C}}}-\overset{O}{\overset{\|}{C}}-O-H \qquad H-\overset{H}{\overset{|}{N}}-\underset{\underset{H}{|}}{\overset{H}{\overset{|}{C}}}-\overset{O}{\overset{\|}{C}}-O-H$$

10. Join the two models together by removing water in the form of two hydrogen atoms and one oxygen atom. Fasten the remaining parts of the two models together. There should be no sticks left over. This process is polymerization.

11. Write the structural formula for the new model.


Hey, there goes my pal Amino . . . so can Tonto be far behind?

12. Join your model to that made by another student. Each time remove a water molecule. *The kind of polymerization in which a small molecule is removed is called CONDENSATION POLYMERIZATION.*

13. Continue to join models with the other students until the class has one giant molecule.

14. Inspect the molecule at the places where the smaller models were joined. What group seem to be repeated? ______________ Draw the structural model of this.

Questions to answer

1. You have seen and used many objects made of polyethylene. The term "poly" means many. Many ethylene molecules get together to form polyethylene. Work out how you think it should be done by using models.

2. Ethylene is called a MONOMER of polyethylene. *A MONOMER is the name of the unit which can be used to make a polymer.* What do you think is the monomer of polypropylene plastics? ______________

____

EXERCISE

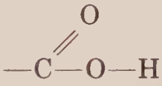
CHEMICAL HIDE AND SEEK

Compounds are classified according to the functional group or groups which are present in their molecular structure. Read Chapter 15 in the Reader-text.

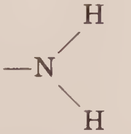
Some classes of compounds and their functional groups are:

ALCOHOL —O—H

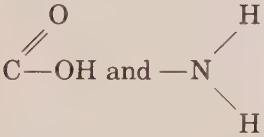
ACID



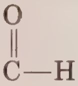
AMINE



AMINO ACID

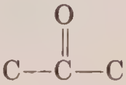


ALDEHYDE



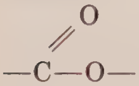
$$\begin{array}{ccc} O & & \\ \| & & \\ C & — & H \end{array}$$

KETONE



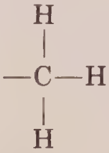
$$\begin{array}{ccccc} & & O & & \\ & & \| & & \\ C & — & C & — & C \end{array}$$

ESTER



ETHER —C—O—C—

METHYL



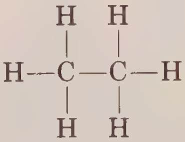
$$\begin{array}{cccc} & & H & \\ & & | & \\ — & C & — & H \\ & | & & \\ & H & & \end{array}$$

Other compounds are classified according to the bonding between carbon atoms in hydrocarbons:
ALKANES C—C ALKENES—C=C—
ALKYNES—C≡C—

The object of Chemical Hide and Seek is to find the hidden clues (functional groups or carbon to carbon bonding) and "call out" by writing the name of the class of compound beside each "formula." The student who gets the most right is the winner.

In this game you are just trying to find functional groups in the formulas given and are not trying to "call out" the correct names of the compounds.

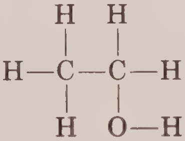
1.



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & — & C & — & C & — & H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

____

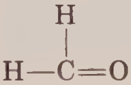
2.



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & — & C & — & C & — & H \\ & & | & & | & & \\ & & H & & O & — & H \end{array}$$

____

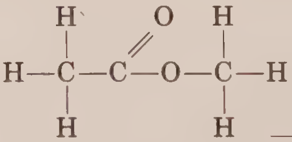
3.



$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & — & C & = & O \end{array}$$

____

4.



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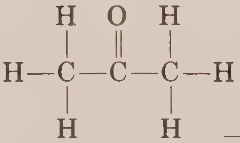
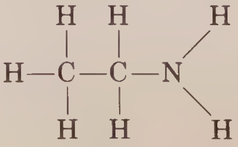
5.



$$\begin{array}{ccccccccc} & & H & & O & & H & & \\ & & | & & \| & & | & & \\ H & — & C & — & C & — & C & — & H \\ & & | & & & & | & & \\ & & H & & & & H & & \end{array}$$

____

6.



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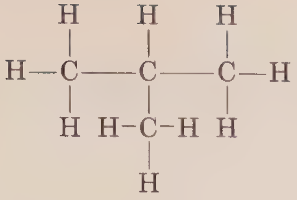
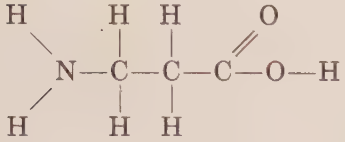
7. 
$H-\underset{H}{\overset{H}{C}}-\underset{H-\underset{H}{C}-H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$

____

8. 

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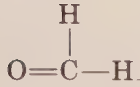
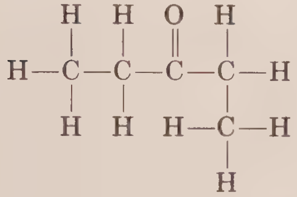
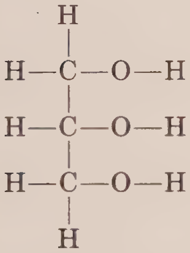
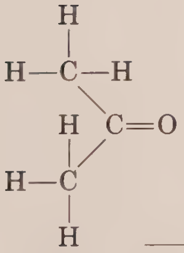
9. 
$O=\overset{H}{C}-H$ ________________

10. 
$H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\overset{O}{\overset{\|}{C}}-\underset{H-\underset{H}{C}-H}{\overset{H}{C}}-H$

____

11. 
$$\begin{array}{rcl} & H & \\ & | & \\ H- & C & -O-H \\ & | & \\ H- & C & -O-H \\ & | & \\ H- & C & -O-H \\ & | & \\ & H & \end{array}$$
____

12. 
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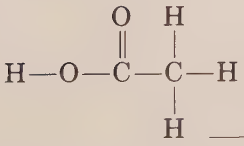
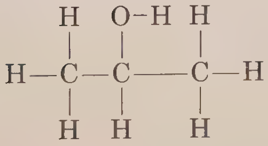
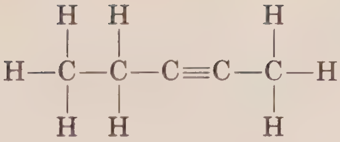
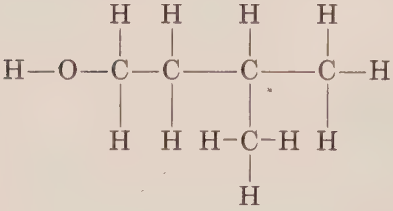
13. 
$H-O-\overset{O}{\overset{\|}{C}}-\underset{H}{\overset{H}{C}}-H$ ________________

14. 
$H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{O-H}{C}}-\underset{H}{\overset{H}{C}}-H$ ________________

15. 
$H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-C\equiv C-\underset{H}{\overset{H}{C}}-H$

____

16. 
$H-O-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H-\underset{H}{C}-H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$

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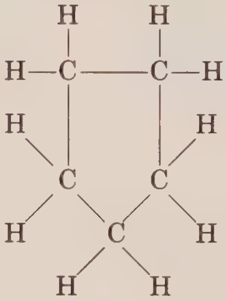
17. 
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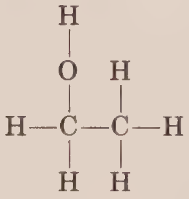
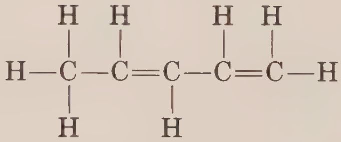
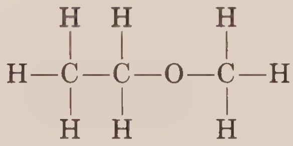
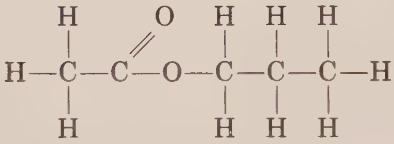
18. 
$H-\underset{H}{\overset{\overset{H}{O}}{C}}-\underset{H}{\overset{H}{C}}-H$ ________________

19. 
$H-\underset{H}{\overset{H}{C}}-\overset{H}{C}=\underset{H}{C}-\overset{H}{C}=\overset{H}{C}-H$

____

20. 
$H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-O-\underset{H}{\overset{H}{C}}-H$

____

21. 

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Chapter 16

GO POWER

Where do you get your go power? Try going without food for a day or two and find out! Do all foods have the same amount of go power? This experiment may give you an answer.

Experiment 16.1

Equipment and Supplies per Lab Station

crucible and cover; Erlenmeyer flask —125 ml
thermometer, —10°C to 110°C; burner; clamp
ring stand; heavy mounting pin; metric ruler
piece of soft string, 1 inch long
3 samples, one of each of three different foods: two solid, one an oil
200 ml of water; 1 or 2 ice cubes

What to do

PART I

1. Obtain two different 0.5 g to 1.0 g samples of solid foods. *Note:* One peanut with the shell and brown covering removed, one-half of a shelled almond, one-fourth of a shelled walnut, each weigh about 1 gram.
2. Place the cover upside down on a crucible. Lay a mounting pin and one of the samples, 1, on the cover. Weigh the crucible, cover, pin, and sample 1. Record the mass in Data Chart 1.
3. Cool about 200 ml of water with ice to about 10 C° *below* room temperature. Place 50 ml of the cooled water in the 125-ml flask.
4. Assemble the apparatus shown in Fig. 73. Make sure that the bottom of the flask is about 2 inches above the weighed crucible and sample.
5. Take the temperature of the water in the flask. Record in Data Chart 2 on page 6/**16**.
6. Loosen the clamp on the ring stand. Turn the flask away from the crucible and sample. Fasten the clamp in the same position as Fig. 73.

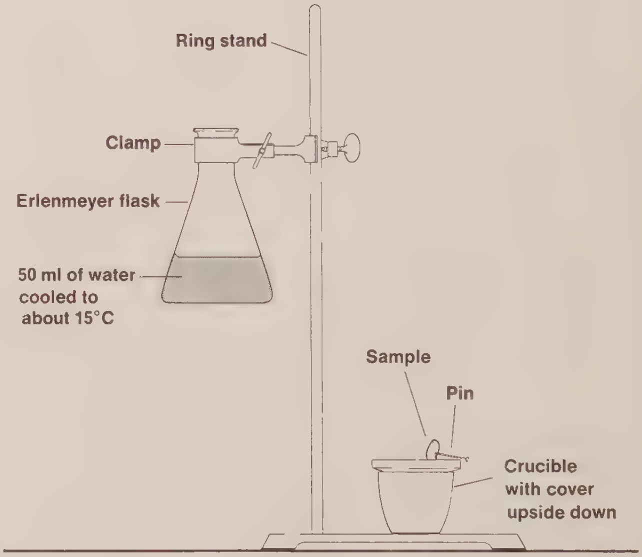
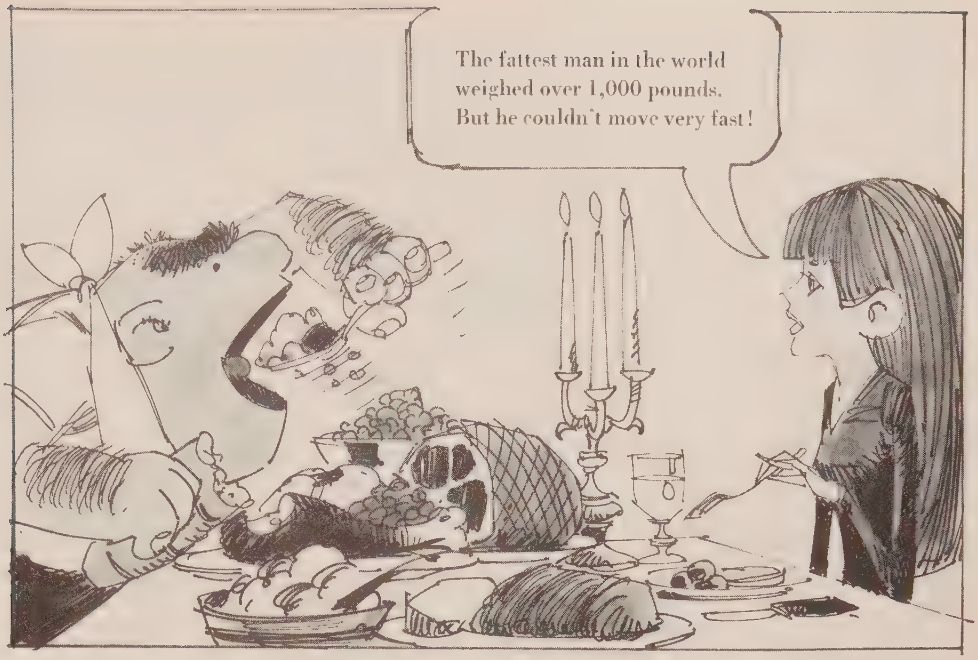

Ring stand
Clamp
Erlenmeyer flask
50 ml of water cooled to about 15°C
Sample
Pin
Crucible with cover upside down

Fig. 73. Set up for PART I. The flask is on the wrong side of the ring stand to allow for adjustment and ignition of sample.

DATA CHART 1

Food sample	Crucible, cover, pin, and sample grams	Crucible, cover, pin, and residue grams	*Mass of sample burned grams
1.			*
2.			*
3. Oil			*

*Subtract the grams of crucible, cover, pin, and *residue* from the grams of crucible, cover, pin, and *sample* at the start.


The fattest man in the world weighed over 1,000 pounds. But he couldn't move very fast!

7. Attach the longer side of the sample securely to the point of the pin. Support the sample with as little as possible of the sample resting on the crucible cover.
8. Make the flame of the burner short but very hot. Heat the top of the sample until it starts to burn. Place a thermometer in the water.
9. Remove the burner. Turn the flask of water *quickly* back into position over the burning sample. Allow the sample to burn for *two* minutes. Stir the water *gently* with the thermometer during the two minutes. Watch the temperature closely.
10. Blow out the flame of the sample. Watch the temperature of the water until it just starts to drop. Record the *highest* temperature observed in Data Chart 2.
11. Cool the crucible, cover, pin, and what remains of the sample (the residue). Weigh the crucible, cover, pin, and residue. Record the mass in Data Chart 1.
12. Wash and dry the crucible cover and pin.
13. Place solid sample 2 on the crucible cover. Weigh the crucible, cover, pin, and sample 2. Record the mass in Data Chart 2.
14. Replace the water in the flask with 50 ml of cooled water. Take the temperature of the water. Record in Data Chart 2.
15. Repeat steps 6 through 11 with sample 2.
16. Wash and dry the crucible cover and pin for use in PART II.

PART II

17. Obtain a sample of oil by placing two drops of oil, about 0.5 g, on the crucible cover. Place about one inch of soft string in the oil.
18. Weigh the crucible, cover, oil, and string. Record the mass in Data Chart 2.

19. Replace the water in the flask with 50 ml of cooled water. Take the temperature of the water. Record in Data Chart 2.
20. Arrange the apparatus as in PART I. Turn the flask of water away from the crucible and sample.
21. Use the point of the pin to pull one end of the oily string to the edge of the crucible cover.
22. Hold a short but hot burner flame at the tip of the string until the string starts to burn. See Fig. 74. Remove the burner. Place a thermometer in the water.
23. Turn the flask of water *quickly* back over the burning sample. Stir the water gently with the thermometer.
24. Allow the oil to burn for two minutes. Place an asbestos pad over the crucible cover to stop the burning. Watch the temperature of the water until it starts to drop. Record the *highest* temperature observed in Data Chart 2.
25. Cool the crucible, cover, and residue. Weigh as in PART I. Record the mass in Data Chart 1.

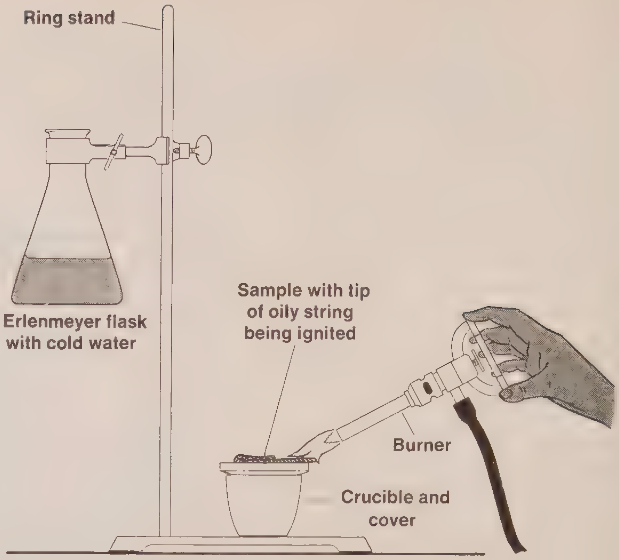

Ring stand
Erlenmeyer flask with cold water
Sample with tip of oily string being ignited
Burner
Crucible and cover

Fig. 74. Set up for PART II. The flask is on the wrong side of the ring stand to allow the ignition of sample.

DATA CHART 2

Food sample	Temperature at the start °C	Temperature after 2 min. °C	*Temperature change C°
1.			*
2.			*
3. Oil			*

*Subtract the temperature at the *start* from the temperature after the two minutes.

Calculations

1. Calculate the mass burned for each sample in Data Chart 1.
2. Calculate the increase in temperature of the water for each sample in Data Chart 2.

Questions to answer

1. Did the same amount of each food burn in the two minutes? ______________ Might the amount of heat given off for the amount of food burned have something to do with the amount of "go power" produced? ______________________________

____

2. If we say "go power" is the amount of heat given off by a substance when it burns, would you think the three foods you tested produced the same amount of "go power?" Explain. ______________________

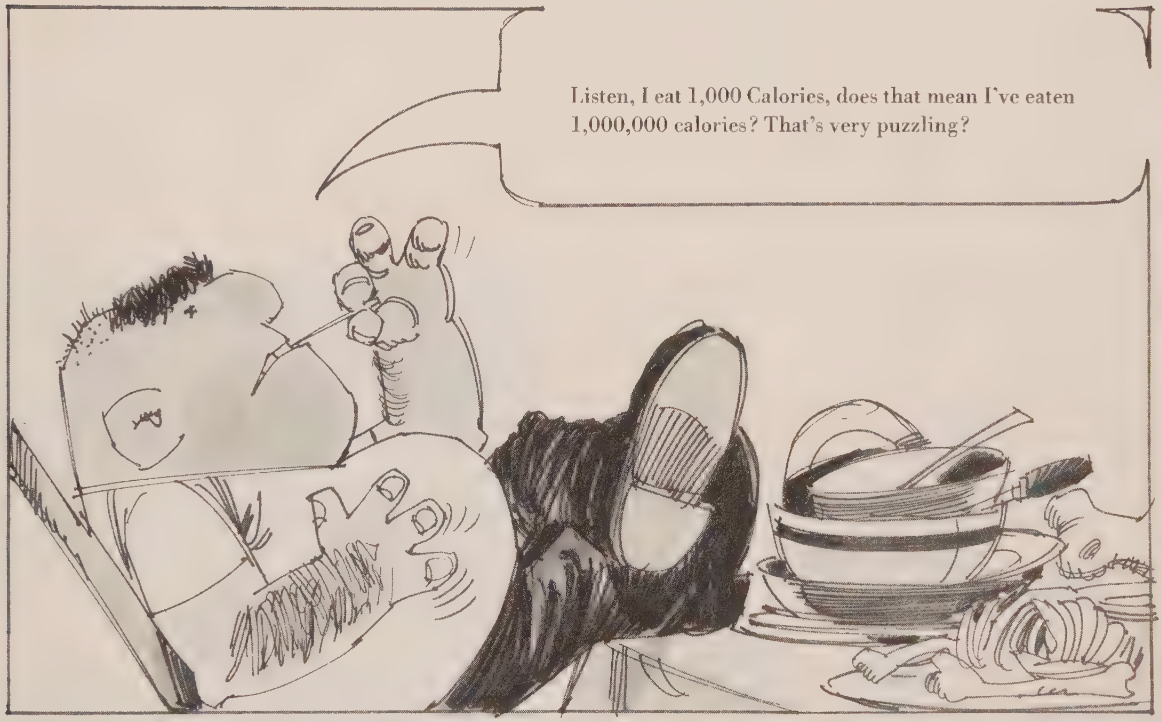
____

____

What is a **Calorie?**

In Experiment 16.1 we were interested in Go Power. Go Power is another term for ENERGY. *ENERGY is the ability to do work.* Heat is a form of energy. The chemist uses a CALORIE as a unit to measure heat. What is a calorie?

People on diets often talk about the calories in food. These calories are large calories (Cal). They are 1000 times larger than the basic unit calorie (cal). We can find the food calories in the samples you burned in Experiment 16.1 by a few calculations.


Listen, I eat 1,000 Calories, does that mean I've eaten 1,000,000 calories? That's very puzzling?

Exercise

You need to know:

1. the mass of the sample burned. Find this out from Data Chart 1 in Experiment 16.1.
2. the temperature change in C° from Data Chart 2.
3. the mass of water heated; 1 ml = 1 g. The mass of water used was 50 g.
4. the definition of a calorie. *A CALORIE is the amount of heat needed to increase the temperature of one gram of pure water one Celsius degree.*
5. the method to calculate Food Calories produced from sample burned.

$$\text{Food Calories} = \frac{\text{g of water warmed} \times \text{temp. increase (C°)} \times 1\ \text{cal/g/C°}}{1000\ \text{cal/Cal}}$$

DATA CHART 1

Sample	g of sample burned	g of water heated	Temp. change (C°)	Food Cal	Food Cal/g
#1					
#2					
#3					

CALCULATIONS

(a) Use the method given on page 6/**17** to find the Food Calories. Record in the Data Chart.

(b) Use the following method to find the Food Calories per gram of the substances burned.

$$\text{Food Calories per gram} = \frac{\text{Calories produced}}{\text{g of sample burned}}$$

Record in the Data Chart.

DATA CHART 2

Sample	Accepted Cal/g	Your results Cal/g	Agreement
#1			
#2			
Oil sample			

Questions to answer

1. Why did you have to find the Calories produced per gram? ________________________________

____

Why aren't the Calories produced all you have to calculate? ________________________________

____

____

2. Which of the samples you tested gives the most "go power?" ________________________________

____

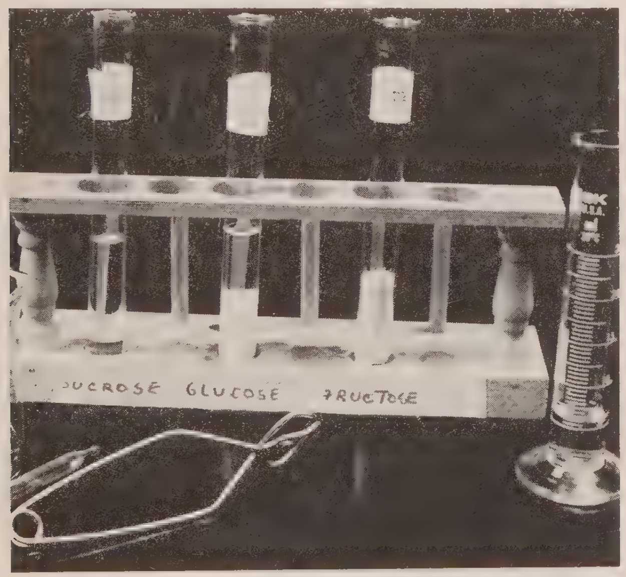
3. Your teacher will help you to find the accepted Calories for each of the foods you tested. How well did your results agree with the accepted Food Calories? ________________________________

Sugars

Familiar sugars are table sugar, corn syrup, and the sweetness of fruits. The chemist knows table sugar as *sucrose,* corn syrup as *glucose,* and the sweetness of fruits as *fructose.*

Discover for yourself likenesses and differences among these sugars.

Fig. 75. Set up in which sucrose is dissolved, glucose has water added, and fructose is without water. *(Clara Aich and George Senty)*



Experiment 16.2

Equipment and Supplies per Lab Station

(2) beakers – 100 or 150 ml; (4) droppers with extra long tips
funnel; graduated cylinder; glass stirring rod
(12) test tubes – 18 mm × 150 mm; balance; burner
filter rack or small ring for ring stand
ring stand with ring and wire gauze
solid stopper to fit the test tubes
test tube holder or clamp; test tube rack
filter paper, fine grade to hold back fine precipitates
samples of sweet food products, solids and liquids
6-8 g fructose; 6-8 g glucose; 6-8 g sucrose; distilled water
15 ml of ethanol; 5-10 ml concentrated sulfuric acid, H_2SO_4
40-50 ml Benedict's solution
5 ml saturated solution barium hydroxide, $Ba(OH)_2$

What to do

PART I

1. Obtain samples of about 8 grams of each of the three sugars: sucrose, glucose, and fructose. Label each.
2. Solubility
    - (a) Label a test tube for each of the sugars.
    - (b) Place five grams of each of the sugars in the separate labeled test tubes. Measure the height of the dry sugar in each test tube. Record in Data Chart 1.
    - (c) Add 5 ml of water to each test tube. Stopper and shake well each in turn. Let stand until any undissolved sugar settles.
    - (d) Measure the height of any undissolved sugar in each test tube. Record in Data Chart 1. *Save these solutions for step 3.*
    - (e) Repeat (a) through (d). Use 1 g of each sugar instead of 5 g. Add 5 ml of ethanol instead of water. Record your observations in Data Chart 1. You may discard these ethanol solutions.

DATA CHART 1

Sample	Water		Ethanol	
	Dry, cm	Undissolved, cm	Dry, cm	Undissolved, cm
Sucrose				
Glucose				
Fructose				

(f) Record in Data Chart 2 which sugar is most soluble, and least soluble, first in water and then in ethanol.

3. Benedict's Test
   (a) Place a separate clean long-tipped dropper in each test tube with the *water* solution of one of the sugars.
   (b) Place 5 ml of Benedict's Solution in a clean empty test tube. Heat the solution just to boiling.
   (c) Add two to five drops, one at a time, of the sucrose solution. Heat for about 10 seconds after each addition of sucrose solution. Record your observations in Data Chart 2.
   (d) Repeat (b), and (c), first with the glucose solution and then with the fructose solution. Record your observations in Data Chart 2.

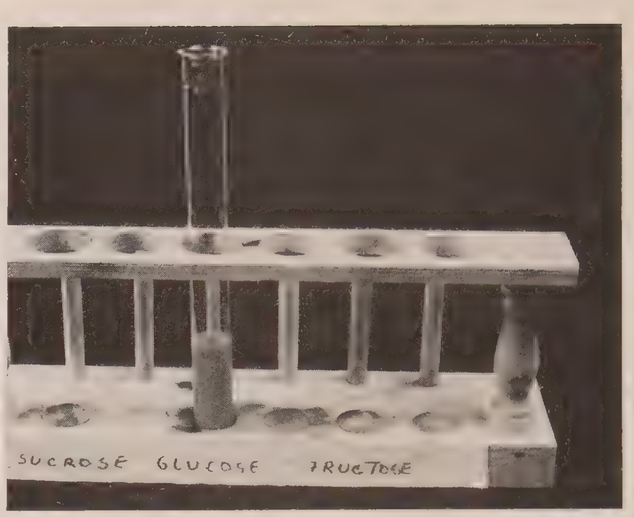


Fig. 76. Benedict's solution is used to indicate the presence of glucose. *(Clara Aich and George Senty)*

DATA CHART 2

Sugar	Solubility in Water	Solubility in Ethanol	Benedict's Test
Sucrose			
Glucose			
Fructose			

4. Unknown Samples
   (a) Obtain samples of two or more common sweet food products. Some may be dry solids. Others may be liquids.
   (b) Test each sample to discover what sugar it contains. Use your imagination to find appropriate ways to use the solubility test and the Benedict's Test.
   (c) Describe your methods. ______________________

   ____

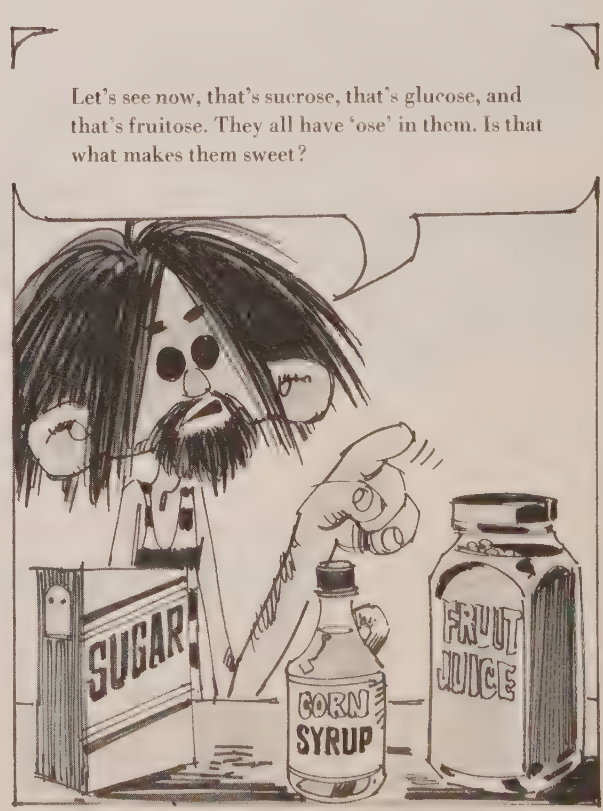
   ____

   ____

   (d) Record in Data Chart 3 what sugar you think may be present in each food tested. Record also your reason for deciding on each sugar.

PART II Can sucrose be changed to other sugars?

1. Use a 150-ml beaker. Dissolve 2 g of sucrose in 20 ml of distilled water. Add 4 drops of concentrated sulfuric acid.
   CAUTION: REMEMBER THAT CONCENTRATED SULFURIC ACID BURNS SKIN AND CLOTHING SERIOUSLY.
2. Heat the solution to boiling. Boil for one minute. Cool to room temperature.
3. Add 5 ml of a saturated solution of barium hydroxide to the cool solution. A precipitate will form.


Let's see now, that's sucrose, that's glucose, and that's fruitose. They all have 'ose' in them. Is that what makes them sweet?
SUGAR
CORN SYRUP
FRUIT JUICE

4. Filter to remove the precipitate.
5. Test the clear filtrate with Benedict's Solution. What do you observe? ______

DATA CHART 3

Food Product	Sugar Present	Reason
#1		
#2		
#3		

Questions to answer

1. Ask your teacher for the accepted solubilities of each of the three sugars in water and ethanol. How do your results compare with the accepted solubilities?
2. What differences did you observe in the reactions of the three sugars with Benedict's Solution? ______
3. In PART II, steps 1 through 4, you prepared a different kind of solution from sucrose. Was the Benedict's Test on this solution positive or negative for simple sugar? ______
4. Look up the structural formula of each of the sugars tested. Write the formula in the space provided below.
   (a) Sucrose
   (b) Glucose
   (c) Fructose
5. What do you think might have happened to the sucrose when it was heated with the sulfuric acid? Explain. ______

ABOUT STARCH

Look—starch shakes out white! How can red turn it blue?

Does starch form a true solution with hot water or does it form a colloidal suspension? Remember that a true *solution* is clear. The particles of solute are molecules or ions much too small to be seen. They are also too small to show the pathway of a beam of light passing through the solution.

A liquid *colloidal suspension* may be evenly cloudy. Large clumps of molecules are distributed throughout the liquid. These large particles scatter light so that the pathway of a beam of light passing through can be seen.

Experiment 16.3

Equipment and Supplies Per Lab Station

(2) beakers—125 ml; crucible; funnel with filter paper
test tube with solid stopper to fit; graduated cylinder—50 ml
balance; burner; flashlight
ring stand with ring and wire gauze
ring stand with small ring or filter stand
soluble starch, 2 grams; 100 ml—iodine test solution

What to do

Test for starch

1. Place a small amount of powdered starch in your crucible. Drop on it one or two drops of iodine test solution.

   (a) What color is the iodine solution? ______________

   (b) What happens when the iodine solution is dropped on the starch? ______________

   ____

2. Test at least three food products for starch.

	Food Product	*Iodine Test*
(a)	______________	______________
(b)	______________	______________
(c)	______________	______________

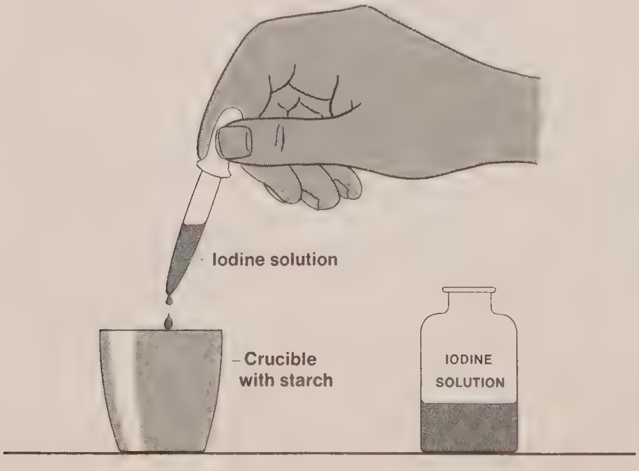


Fig. 78 (A, B, C). Solubility of starch.

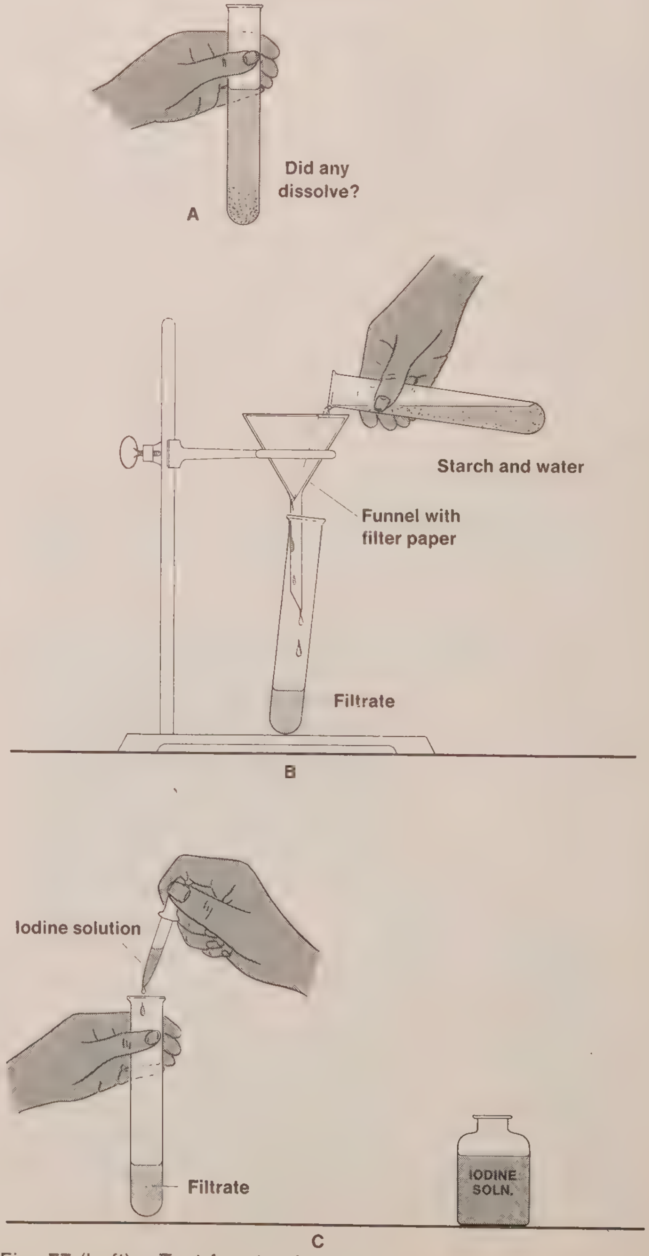


Fig. 77 (Left). Test for starch.

Solubility of Starch

3. Place about 1 ml of starch in a test tube. Add 10 ml of water. Stopper and shake vigorously. Let stand one or two minutes. Is the volume of the starch about the same or less than it was before the water was added? ______________________
4. See Fig. 78. Filter a small amount of the starch and water mixture into a test tube. Add two or three drops of the iodine solution to the clear filtrate. Hold the test tube over a white paper. What do you observe? ______________________

   ____

   Has any starch dissolved in the cold water? __________

   ____
5. Place 1 g of starch in a small beaker. Add 2 to 5 ml of water. Stir to make a paste. See Fig. 79.
6. Heat 95 ml of water to boiling in a 150-ml beaker. Turn out the fire. Stir in the starch paste. Describe what happens.

   ____

   ____
7. Cool the liquid until you can hold the beaker. Pour about 5 ml of the liquid into a test tube. Add two to three drops of the iodine test solution. Hold the test tube over a white background. What do you observe?

   ____
8. Place a beaker of tap water or of a solution of sodium chloride on the left side of the beaker of starch liquid. Direct the beam of a flashlight so that it will pass first through the water or salt solution and then through the starch liquid. Ask someone to hold the flashlight for you.
   Note: This should be done away from *direct* outside light or *direct* room lights, if possible. It need not be done in darkness.
9. (a) Mark what you see when the light is passed through the two beakers.
   (b) Describe what you observe. ______________________

   ____

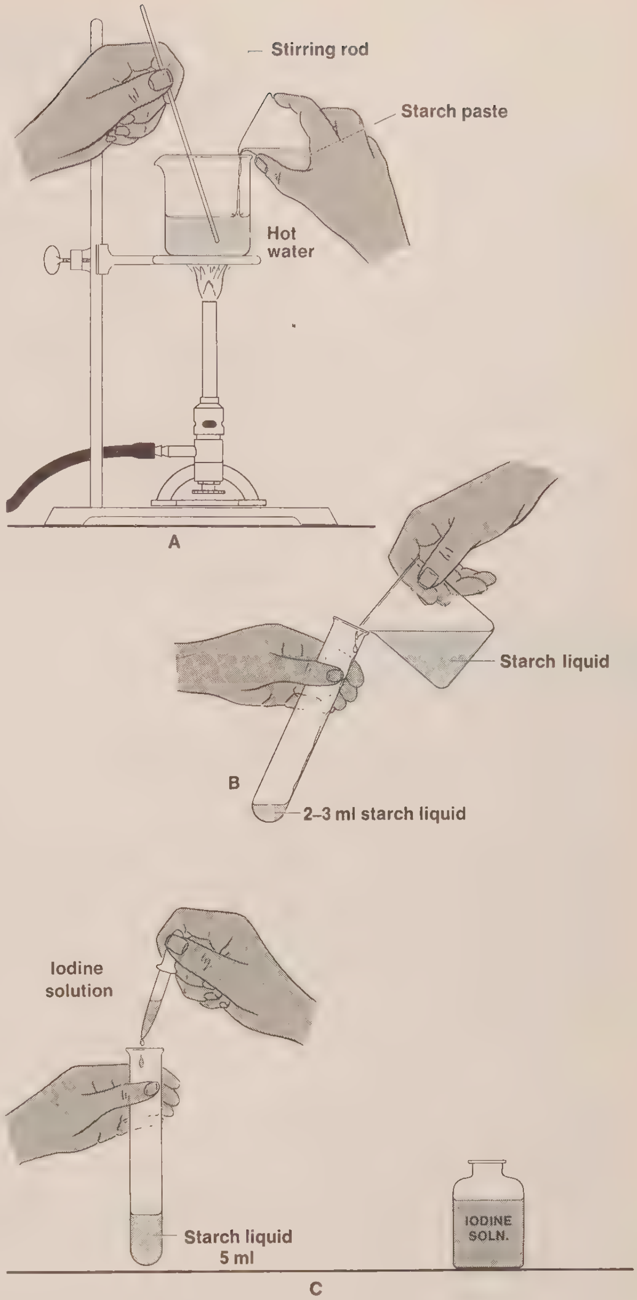


Fig. 79. Solubility of starch.

Questions to answer

1. Describe the test for starch. ______________________

   ____
2. Do you think the starch dissolved in the hot water?

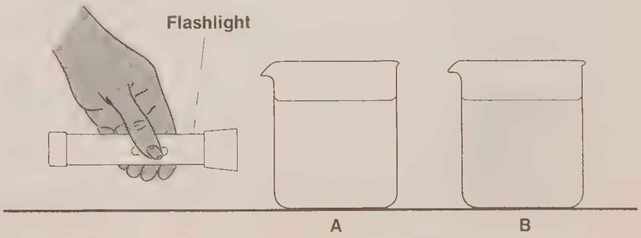
   ____

   If not, what was formed? ______________________

   ____
3. What is cornstarch chocolate pudding? ______________________

   ____

Fig. 80. The diffusion of light by particles is known as the TYNDALL EFFECT.



ABOUT FAT

Fat—who wants it? But we all need some in our food to stay healthy.

Many foods contain fat. Sometimes it is easy to see, as in meat; sometimes it doesn't show, as in egg yolk.

Oils are fats. Vegetable oils, like salad oil, can be changed into solid fats. This is how oleomargarine is made. The process is called HYDROGENATION.

HYDROGENATION is the addition of hydrogen to a substance. Oils become solid fats when some of the molecules in the liquid oil combine with additional hydrogen. Hydrogen can be added to oils in a process using finely divided nickel as a catalyst.

A CATALYST is a substance which increases the rate of a chemical reaction without being used up itself.

What are some special properties by which we can tell fats from other substances? Find out for yourself.

Experiment 16.4

Equipment and Supplies per Lab Station

acetone, 3 ml; benzene, 3 ml; ethanol, 3 ml
distilled water; water, boiling, 200 ml
red and blue litmus paper; hydrochloric acid, dilute, 3 ml
sodium hydroxide, dilute, 3 ml; sodium carbonate solution, 0.5%, in dropper bottle; albumin solution, 5 ml
beaker —250 ml; medicine dropper; (10) test tubes —18 mm × 150 mm
solid stopper to fit the test tubes; test tube rack
(6) small pieces of cloth; vegetable oil, 5 ml

What to do

PART I TESTING FAT

1. Obtain about 5 ml of a vegetable oil.
2. *Spot test:* Put one drop of oil on a piece of notebook paper. What happens? ____________________
3. *Litmus test:* Dampen a strip each of red and blue litmus paper with water. Lay the strips on a glass plate. Put a drop of oil on each. What happens:
   (a) with the red litmus? ____________________
   (b) with the blue litmus? ____________________
4. *Solubility:*
   (a) Place about 3 ml of each of the following solvents into separate *labeled* test tubes.
   Water Dilute base Acetone
   Dilute acid Ethanol Benzene
   CAUTION: MAKE SURE THAT ALL FIRES ARE OUT AS LONG AS ANY ACETONE OR BENZENE IS AROUND. KEEP ETHANOL AWAY FROM AN OPEN FIRE.
   (b) Add five drops of the oil to each of the solvents in its test tube.
   (c) Stopper and shake the test tubes, one at a time. Let stand one to two minutes. Record your observations in the Data Chart.
   (d) Pour the tests with acetone and benzene into the sink. Wash away with plenty of water.
   (e) Transfer the tests with water, acid, base, and ethanol to a beaker of hot water. Get the hot water from a hot water tap or from your teacher. (Remember: NO FIRES!)
   (f) Keep the tests warm three to five minutes. Examine the tests. Record your observations in the Data Chart.
   (g) Which are the best solvents for fat? ____________________
   ____
5. Place one drop of the oil on each of several small pieces of cloth. Put one oily piece of cloth into each of the solvents in separate test tubes. Which of the solvents removed the spot most quickly and completely? ____________________

DATA CHART

Solvent	Cold	Hot
Water		
Dilute acid		
Dilute base		
Ethanol		
Acetone		not done
Benzene		not done

PART II EMULSIFICATION

EMULSIFICATION prevents a fat from forming large drops and separating from a liquid. Fats can be emulsified by breaking them into extremely small droplets. Sometimes, also, a substance called an emulsifying agent is added. Milk is homogenized by breaking the cream into tiny droplets.

HOMOGENIZED means made all alike. "Homo" means alike. The milk is made alike by blending the fat (cream) in the remaining liquid so that the fat does not separate and come to the top of the milk.

Do you think fat is heavier or lighter than water?

____

____

What to do

1. Shake a drop of oil with 2 ml of water in a test tube.
   (a) Do small drops form? ____________________
   (b) Do the drops remain small when the test tube stands for one or two minutes? ____________
2. Add three drops of 0.5% sodium carbonate solution to 5 ml of water in a test tube. Add one to two drops of the oil. Shake well. Let stand. What do you observe? ______________________________
   ____
3. Shake a drop of the oil with 5 ml of albumin solution. Let stand. What do you observe? __________
   ____

Questions to answer

1. Describe a simple test to tell whether or not a substance is a fat. ____________________

____

____

____

2. Is fat acid or base? ____________________

3. Which solvent (or solvents) worked best when hot?

____

____

4. Soap is basic in reaction. Why is soapy water for cleaning dishes or clothes usually used warm? ______

____

5. Why do you shake oil and vinegar salad dressing just before you put it on your salad? ____________________

6. What do you think was the job of the sodium carbonate solution and also of the albumin solution?

____

Which worked better? ____________________

Imagine that! You can use a catalyst to make things happen faster... without using the catalyst up! How can that happen? It's causing all kinds of trouble in that reaction, and nothing happens to it! Wow!

Extra for experts

1. Ask someone who knows how to make mayonnaise how the oil and vinegar are made to stay mixed smooth and creamy. ____________________

____

2. Look up the structural formula for a fat. Write the formula here.

3. Compare the formula of the fat with the formulas of the sugars.

(a) Are they composed of the same chemical elements? ____________________

____

(b) How are they alike? ____________________

____

(c) How are they different? ____________________

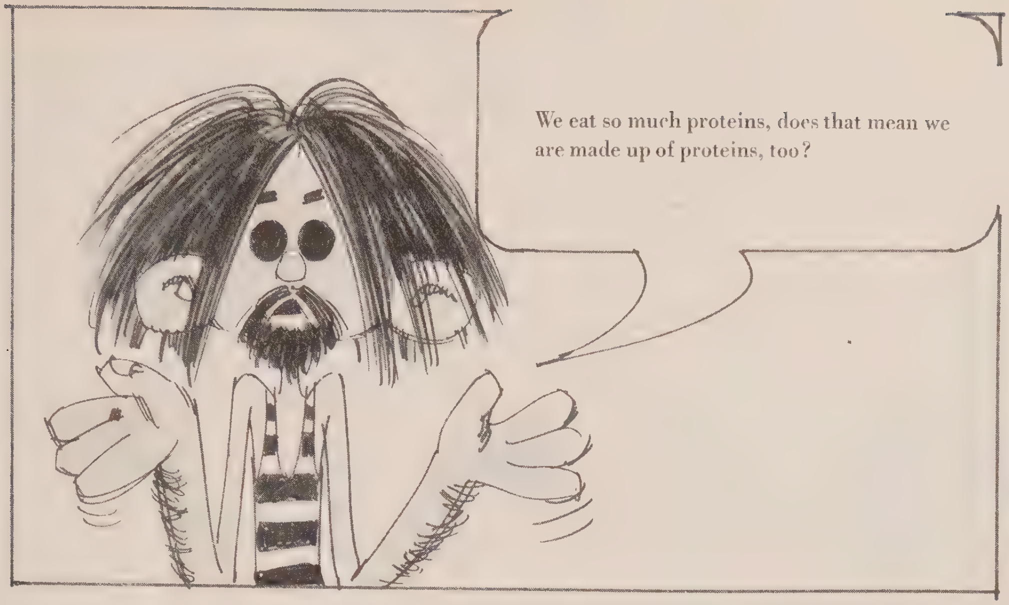
____

____

____

____

____


We eat so much proteins, does that mean we are made up of proteins, too?

ABOUT PROTEINS

Protein food is something you can't live without. Fortunately, there are many protein foods; and they taste good too—meat, fish, gelatin, eggs, cheese, milk, and such vegetables as beans, peas, and nuts.

What chemical elements can be found in protein? How can you test for protein?

Experiment 16.5

Equipment and Supplies per Lab Station

PART I
test tube —18 mm × 150 mm; burner; ring stand; buret clamp
copper wire, fine, 22 gauge, $1\frac{1}{2}$ inches
filter paper; red litmus
lead acetate solution in a small dropper bottle
dry gelatin, powdered, about 1 gram
PART II
medicine dropper; (6) test tubes —18 mm × 150 mm
burner; test tube holder
distilled water; ammonium hydroxide, 0.5 *M*
hydrochloric acid, 0.5 *M*; nitric acid, concentrated
sodium hydroxide, 10%, 5 ml; copper sulfate, 0.5%, 1 ml
fresh egg white, approximately 5 ml

What to do

PART I CHEMICAL ELEMENTS IN PROTEIN

1. Place 1 g of dry gelatin in a test tube.
2. Assemble the apparatus as shown in Fig. 81, step by step.

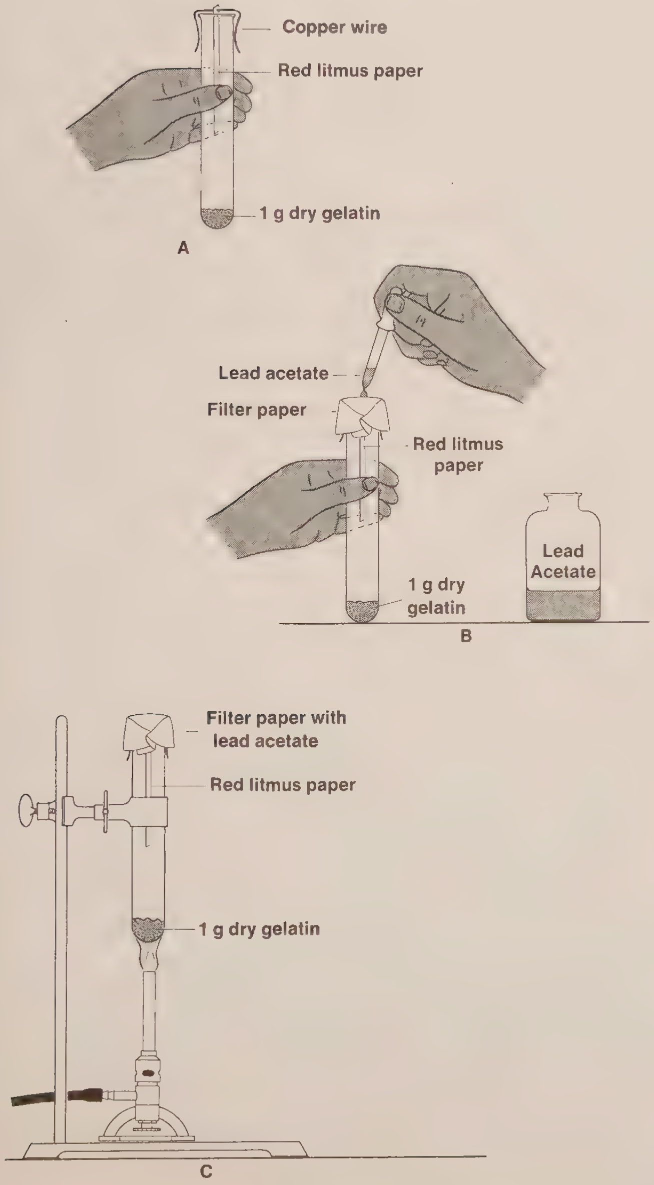


Fig 81. Test for chemical elements in proteins.

3. Heat the bottom end of the test tube gently for 5 to 8 minutes. Record your observations.
   (a) What appears on the inside of the upper part of the test tube? ______________________

   ____

   ____

   What two chemical elements are combined to make this substance? ______________________

   ____

   (b) What change did you see in the litmus paper?

   ____

   A basic reaction in the litmus paper in this experiment would show the presence of nitrogen as ammonia, NH_3. Did your test show the presence of nitrogen? ______________________
   (c) What change takes place in the lead acetate on the filter paper? ______________________

   ____

   ____

   The darkening of lead acetate is a test for sulfur. Did you find sulfur in the gelatin? ______________
   (d) What is left in the test tube at the end of the heating? ______________________
   What chemical element is this most likely to be?

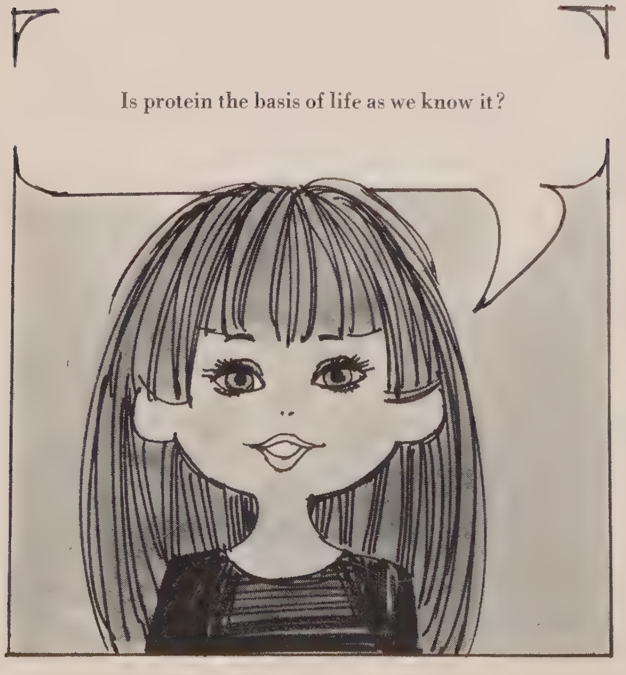
   ____

   ____

4. Compare the chemical elements found in gelatin, a protein, with those found in sugars and fat.

Gelatin	Sugars and Fat

   (a) Which elements are found in all three, protein, sugars, and fat? ______________________
   (b) Which element did you find in the protein but not in the sugars and fat? ______________________


Is protein the basis of life as we know it?

PART II TESTS FOR PROTEIN

1. Place 5 ml of distilled water in each of three numbered test tubes. Add to each about 1 ml of fresh egg white. Stopper each test tube and shake well. (Note: Use distilled water. Tap water contains salts which can interfere in the tests.)

   What do you observe? ______________________

2. Heat the solution in the No. 1 test tube gently.

   Describe any change you observe. ________________
3. Add a few drops of 0.5 *M* hydrochloric acid solution to test tube No. 2. What did you observe? __________

   ____
4. The Xanthoproteic Reaction
   (a) Add about 1 ml of concentrated nitric acid from a dropper bottle to test tube No. 3.

   What do you see happen? ________________
   (b) Heat the solution gently. What changes occur?

   ____

   ____

   ____
   (c) Cool the solution. Add about 5 ml of dilute ammonium hydroxide solution. What further changes do you see? ______________________

   ____
5. The Buret Test
   (a) Place about 0.5 ml of egg white in a test tube. Add 3 ml of water. Then add 4 ml of 10% sodium hydroxide solution. Stopper the test tube and shake well.
   (b) Add 0.5% copper sulfate solution, *drop by drop.* Shake the test tube from side to side to mix after each drop. Watch for a purplish-violet or pinkish-violet color. Do not add too much copper sulfate solution or a precipitate may spoil the test.

   Describe the color changes you observed. _______

   ____

   ____
6. Which is the simplest one of the tests for protein which you did? ______________________

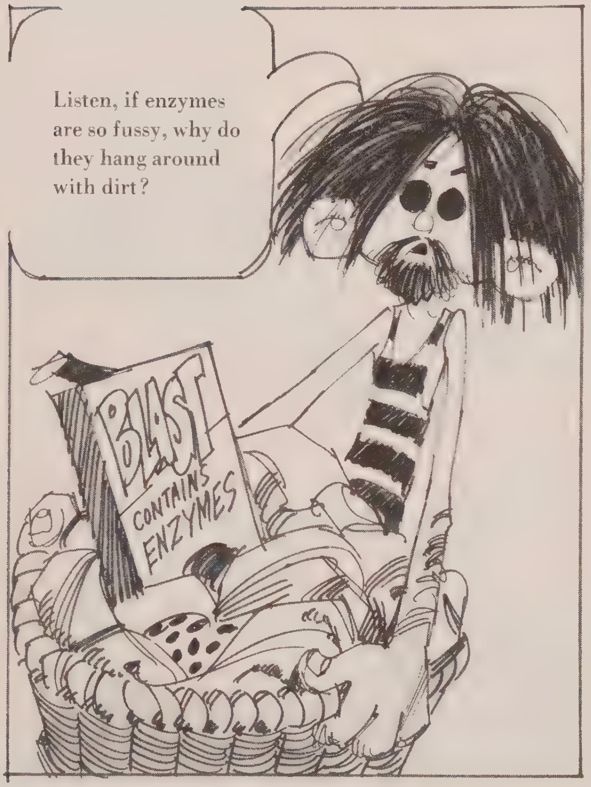
There goes my pal amino!

Questions to answer

1. You learned in Chapter 15 that protein is formed from amino acids. The amino acids differ from each other in structure, but the molecule of each one of them contains an amino group $-NH_2$, and a carboxyl group, $-COOH$. Amino acids join to form proteins by means of the peptide linkage between the amino group of one amino acid and the carboxyl group of another. A molecule of water splits out. Protein molecules are very large and complex, but the peptide linkage between amino acids can be found in them. Write a structural formula in the space at the right showing a peptide linkage between a carboxyl group and an amino group.
2. Why do you think nitrogen is important in the structure of proteins? ______________________

   ____

Enzymes Are Fussy


Listen, if enzymes are so fussy, why do they hang around with dirt?
BLAST
CONTAINS ENZYMES

Your food would do you little good without enzymes to break it down into products your body can use. *ENZYMES are proteins which act as catalysts.* They regulate the rate of chemical reactions in living tissues. They may bond temporarily with the reactants but are not used up in the reactions.

The first of the digestive enzymes is in your mouth. It is called ptyalin or salivary amylase. *AMYLASES are starch-splitting enzymes.* They are found in other parts of the digestive tract as well as in the mouth.

Enzymes are fussy about two conditions, pH and temperature. There is a pH and temperature at which each works best. You will find out about the best pH for salivary amylase. Its favored temperature is 38°C.

The splitting of starch to sugar by salivary amylase can be followed by two tests you have already used.

The Iodine Test

blue-black color -- shows starch
red color ------------ shows dextrins of partly digested starch
no color ------------- shows other dextrins as starch is further digested. Finally, maltose sugar is produced.

Benedict's Test

no change or slight change --- negative test, no simple sugar
yellow precipitate ---------------- positive ++
orange precipitate --------------- positive +++
brick red precipitate ------------ positive ++++

In this experiment you will test the amylase activity of your own saliva.

Experiment 16.6

Equipment and Supplies per Lab Station

beaker —150 ml; beaker —250 ml; medicine dropper; funnel
cheesecloth, 3 inch square, three to four thicknesses
(6) test tubes —18 mm × 150 mm; thermometer
burner; filter stand or ring stand with a small ring
ring stand with ring and wire gauze; test tube holder
paraffin
Benedict's Test Solution, 20-25 ml
Iodine Test Solution, in a dropper bottle
Starch solution, 1%, 10 ml; Universal Indicator paper
saliva collected by each student for his own experiment
(Each student should collect saliva, even though he is working with a partner.)
porcelain spot plate or a glass plate set on white paper ruled in one inch squares

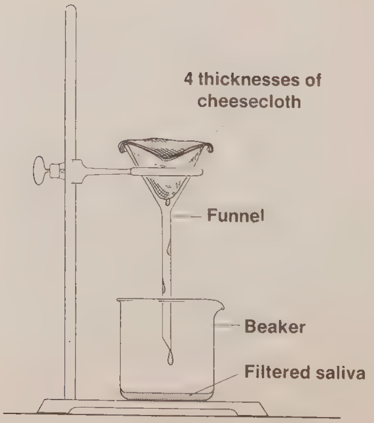


Fig. 82.

What to do

1. Collect about 20 ml of saliva in a small beaker. Your teacher will explain how. Chewing a small piece of paraffin, NOT GUM, will make this easier.

   Why should gum not be used? ______________________________

   ____

2. Filter the saliva through three or four thicknesses of cheesecloth. See Fig. 82.

3. Test the pH of your saliva with Universal Indicator paper. Record. Record also the pH of the saliva of three other students.

   (a) Your saliva – pH ______________________________

   (b) Student 1 – pH ______________________________

   (c) Student 2 – pH ______________________________

   (d) Student 3 – pH ______________________________

4. Fill a 250-ml beaker about half full of tap water. Warm the water to 40°C. Be careful that the temperature *does not go above 40°C.*

5. Place 10 ml of 1% starch solution in a 150-ml beaker. Warm the starch solution to 38°C.

6. Set the beaker with the starch solution into the beaker of water which is now 38°C to 40°C. See Fig. 83.

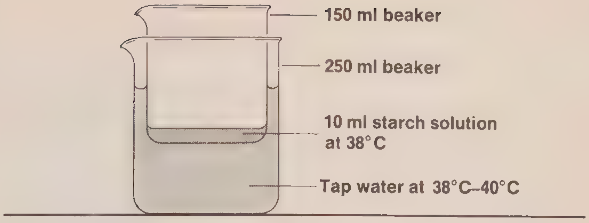


Fig 83.

7. Place one drop of the starch solution in a depression on a porcelain test plate. Add one drop of iodine test solution. The iodine will become blue-black as it did in Experiment 16.3. This test will serve as a control. See Fig. 84.

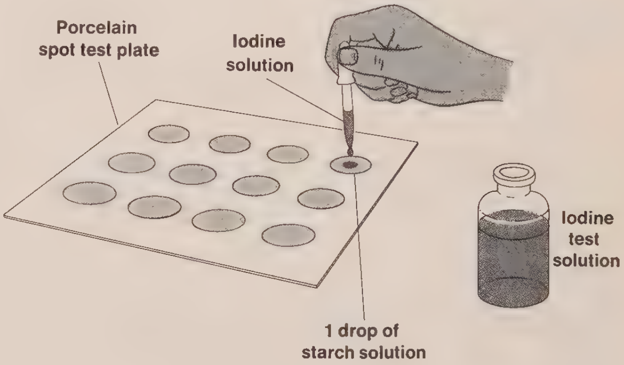
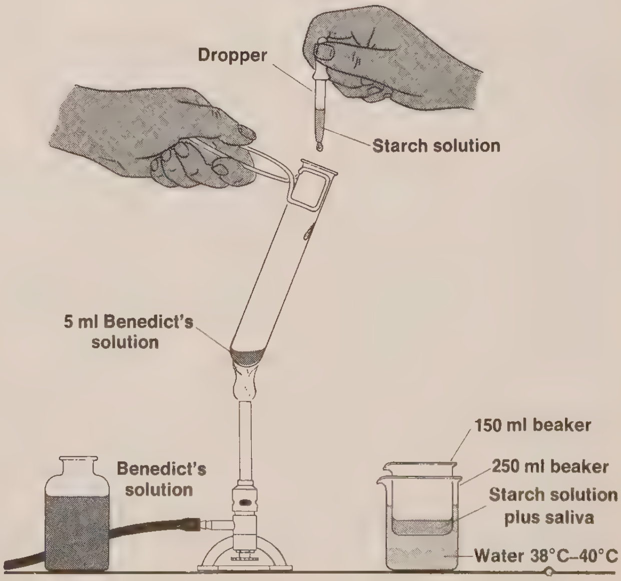


Fig 84.

8. Test the starch solution with the Benedict's Solution as you did in Experiment 16.1. Keep this test also as a control. See Fig. 85. Is the test negative or positive for simple sugar? ______________________________

Fig 85.



9. Add 5 to 10 drops of saliva to the warm starch solution. From now until the end of the experiment *keep the temperature of the starch-saliva solution constant at 38°C.*

10. Choose a time schedule for testing the progress of the digestion of the starch. Different groups of students should choose different time schedules. Record results from both schedules.

    TIME INTERVALS
    Group I——5 minutes, 10 minutes, 15 minutes
    Group II——3 minutes, 9 minutes, 27 minutes

11. At the time intervals chosen, test the starch-saliva solution, first with the iodine and then with the Benedict's Solution. Record the results of all tests in both groups in the Data Chart.
    Iodine test——record the color
    Benedict's test——record: negative, ++, +++, ++++

DATA CHART

Minutes of "digestion"	Iodine Test Color	negative, ++, +++, ++++ Benedict's Test
Group I 5		
10		
15		
Group II 3		
9		
27		

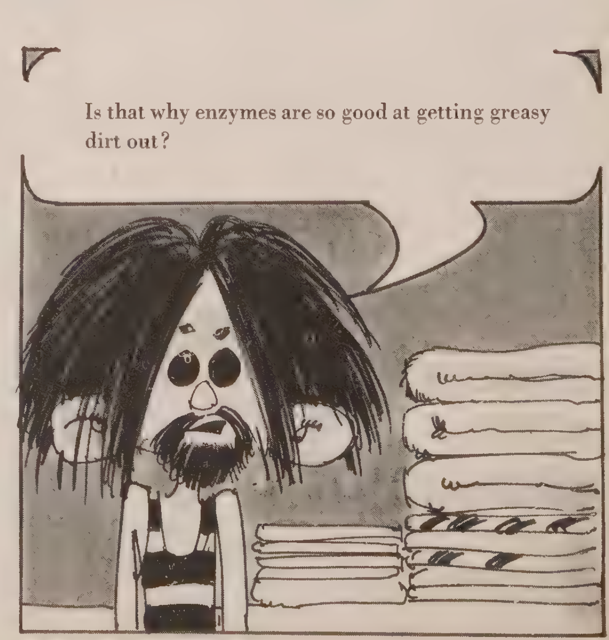
Questions to answer

1. What differences, if any, did you find in the pH of the saliva of different students?
2. Did the pH of the saliva make any difference in the results of the amylase activity tests? Explain.
3. How many minutes was it before the starch-saliva solution produced a *red* color with the iodine?

   Did all students find the same number of minutes?
4. How many minutes was it before the starch-saliva solution produced *no* color with the iodine?
   Did all students find the same number of minutes?
5. Did the results of the Benedict Test change at the same times as the iodine test? Explain.
6. On the basis of your results, what would be the best pH for salivary amylase activity, that is, what pH would produce sugar most quickly?
7. What difference do you think it would make:
   (a) if the digestion were done with ice-cold starch-saliva solution?

   (b) if the digestion were done at 100°C?

Note: You may need to look for these answers in a reference book.


Is that why enzymes are so good at getting greasy dirt out?

Laboratory Experiments in

ACTION CHEMISTRY

UNIT VII

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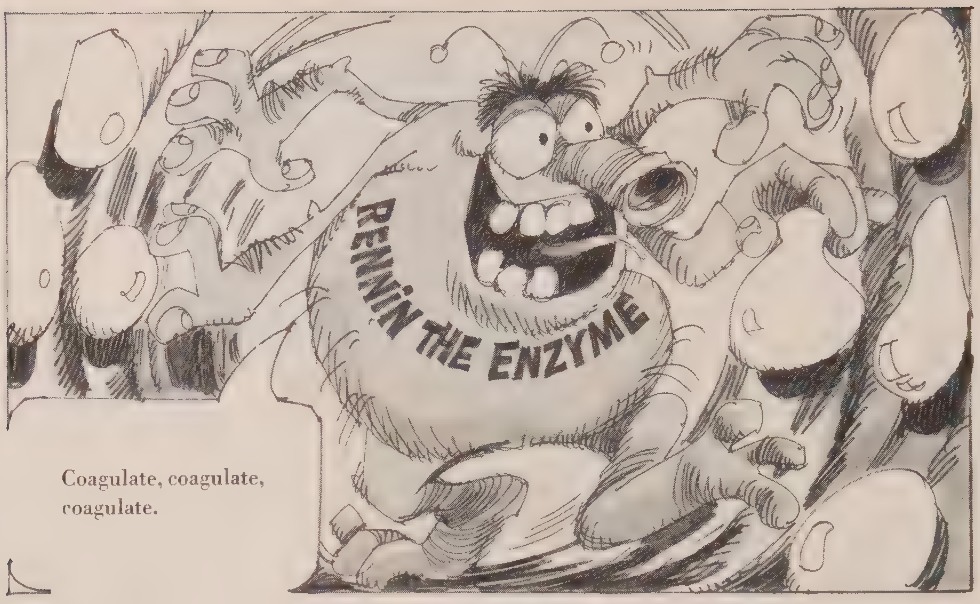
Chapter 17



Coagulate, coagulate, coagulate.

Rennin, An Enzyme

In Experiment 16.6 you discovered an enzyme in your mouth which causes starch to change to simple sugar. Each different enzyme has a highly specialized job. In your digestive tract, some enzymes act on starches and sugars, some on proteins, and some on fats. For each enzyme there is a pH and a temperature at which it acts best.

Rennin is an enzyme found in the inner lining of the stomach of animals which take in milk as food. This enzyme makes the proteins in milk collect together or COAGULATE. The proteins in milk are called casein proteins.

In this experiment you will learn something about the effect of heat on the ability of rennin to coagulate the milk proteins (casein).

Experiment 17.1

Equipment and Supplies per Lab Station
(2) beakers—100 or 150 ml
graduated cylinder; mortar and pestle; (4) test tubes
thermometer; burner; ring stand
timer; wire gauze; junket tablet (rennin); milk

What to do

1. Crush a rennin tablet (junket tablet).
2. Put the crushed rennin into 25 ml of water in a small beaker (100 or 150 ml). Stir to make a fine suspension. Set it aside for later use. Label the Beaker: 1.

3. Repeat steps 1 and 2 with another rennin tablet. Boil the suspension for two to three minutes. Set it aside to cool. Label the Beaker: 2.
4. Pour 10 ml of lukewarm milk (about 35° – 40°C) into each of two test tubes. Label the first Test Tube: 1. Label the second Test Tube: 2.
5. Pour 3 ml of the rennin suspension from Beaker No. 1 into milk sample No. 1.
6. Obtain an empty test tube. Pour the enzyme-milk mixture back and forth between the two test tubes twice. Set Tube No. 1 with the mixture in it aside for eight minutes.
7. Repeat steps 5 and 6 using the milk sample in Test Tube No. 2 and the boiled rennin suspension from Beaker No. 2. 8. Inspect Test Tube No. 1. Hold it over the sink and tip the tube upside down. Write down all of your observations.
8. Repeat step 7 with Test Tube No. 2.
9. Compare the two sets of observations. How can you account for the differences in observations?

   ____

   ____

DATA CHART

	Observations
Test tube No. 1 + Beaker No. 1	
Test tube No. 2 + Beaker No. 2	

Questions to answer

1. If raw pineapple is added to lemon gelatine, the gelatine will not form a gel. If you use cooked pineapple a gel will form. How can you account for this difference? ____________________________

   ____

2. Vegetables are slightly cooked before being frozen. From what you have learned in this experiment can you explain why this is done?

   ____

   ____

   ____

   ____

THE CHEMISTRY OF CANDY-MAKING

Many candies are made from the INVERSION (hydrolysis) of cane sugar. You studied INVERSION in Experiment 16.2. Making these candies involves the control of the growth of crystals from a supersaturated solution of sugars.

In this experiment you will investigate how crystal size can be controlled. You and your partner will be assigned one method to use. Then you can compare your results with the results obtained by your classmates using other methods.

Look at those crystals grow!

Experiment 17.2

Equipment and Supplies per Lab Station

(2) beakers—250 ml; (4) evaporating dishes
thermometer; asbestos pad; beaker tongs; burner
magnifying lens; metal spoon; ring stand with ring
thermometer clamp; wire gauze
cane sugar (sucrose); water
one of the following:
cream of tartar; glycerin; light corn syrup; vinegar

What to do

PART I

1. Place 50 grams of cane sugar and 25 ml of water in a 250 – ml beaker. Stir the mixture, but avoid getting sugar crystals along the sides of the beaker.
2. Set up a ring stand with a ring and asbestos wire gauze, a burner, and a thermometer clamp. See Fig. 86.
3. Set the beaker on the wire gauze. Adjust the thermometer so the bulb is in the sugar and water mixture. Do not let the bulb touch the bottom of the beaker.
4. Boil the mixture over a low flame until 113°C – 115°C is reached. BE CAREFUL! HOT SYRUP CAN CAUSE SERIOUS BURNS IF YOU GET ANY ON YOUR SKIN!
5. Hold the hot beaker with beaker tongs, carefully pour equal portions into each of four evaporating dishes. See Fig. 87.
6. Beat the first portion at once. Beat the other three after they have cooled to 70°C, 60°C, and 40°C respectively.
7. Use a magnifying glass or a low power microscope to examine the size of the crystals formed.
   (a) Which portion seemed to have the smallest crystals? ____________________
   (b) Which portion seemed to have the softest and most creamy texture? ____________________
   ____
   (c) If crystals are to be beaten out of the solution, which temperature appears to be the best one? ____________________
   ____

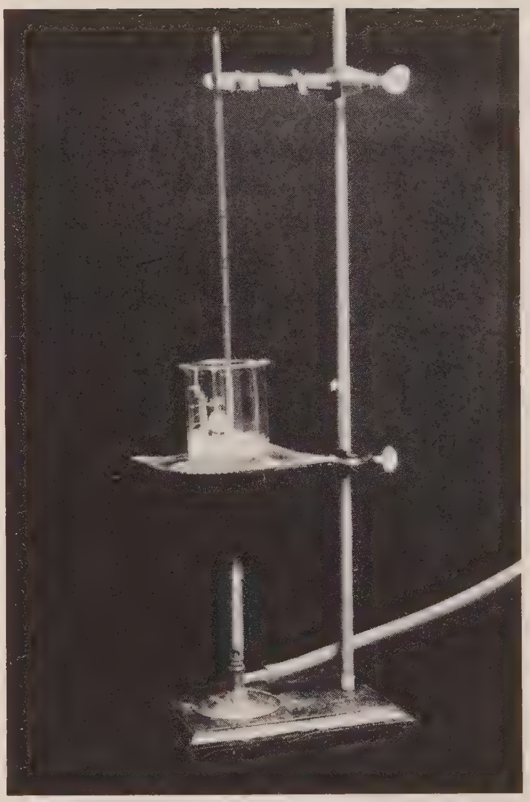


Fig. 86. *(Clara Aich and George Senty)*

PART II

1. Check the assignment given you. Then obtain the materials required by that assignment.

Assignment Number	MATERIALS		
	Water	Cane sugar	Other
1	25 ml	50 g	------
2	25 ml	50 g	0.1 g cream of tartar
3	25 ml	50 g	4 ml vinegar
4	25 ml	50 g	10 ml light corn syrup
5	25 ml	50 g	3 ml glycerine

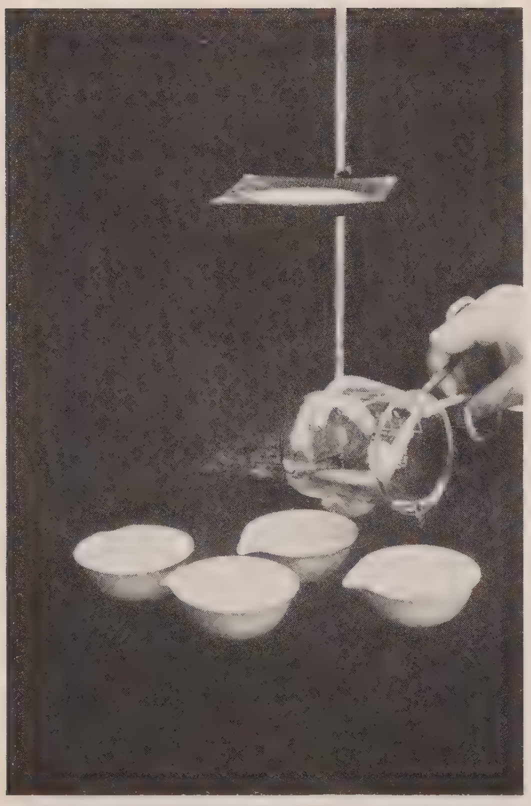


Fig. 87. *(Clara Aich and George Senty)*

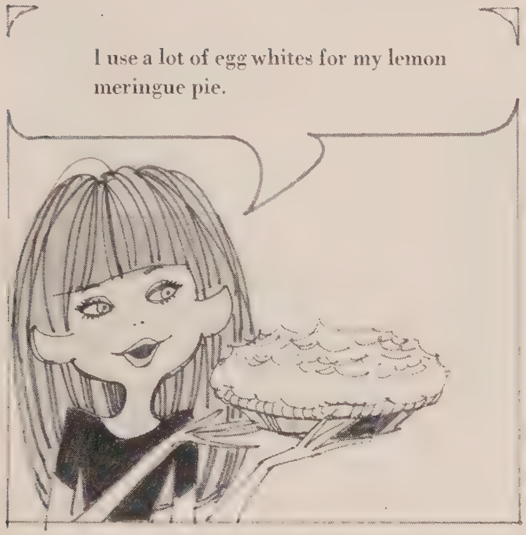
2. Place the assigned materials in a 250-ml beaker. Carefully stir the mixture. Avoid getting sugar crystals along the sides of the beaker.
3. Use the apparatus set up as shown in Fig. 86. Gently heat the substances to 113°C – 115°C.
4. Transfer the beaker to an asbestos pad on the desk with beaker tongs. DO NOT STIR THE MIXTURE OR REMOVE THE THERMOMETER.
5. When the material has cooled to 60°C, remove the thermometer. Beat the mixture until a soft, creamy texture is obtained. (The appearance of the mixture will suddenly seem to change. This is the signal to stop beating.)
6. When the solid has cooled, feel a small portion of it between your fingers. Can you feel hard crystals or is the texture soft and creamy?
7. Compare the textures of the other samples. Which one(s) seem to have a creamy, noncrystalline appearance and feel?
8. What substances can you add to the sugar and water to prevent the growth of large crystals?

Questions to answer

1. Why must you be careful not to get crystals left around the side of a pan when you make a boiled sugar candy?
2. Compare the recipes for making candies. What materials are used to retard crystal growth? What procedures must be used to prevent crystal growth?
3. Write the equation for the hydrolysis of sugar (sucrose).

COAGULATION OF PROTEINS

Proteins collect together or COAGULATE during the preparation of some foods. What are some ways by which the proteins may be coagulated? Find the answer in this experiment.


I use a lot of egg whites for my lemon meringue pie.

Experiment 17.3

Equipment and Supplies per Lab Station

PART I
beaker—250 ml; watch glass; burner; forceps
ring stand with ring; wire gauze; egg white
PART II
dropper; graduated cylinder—10 or 25 ml
test tube; test tube rack
dilute hydrochloric acid, HCl; skim milk (non-modified type)
PART III
graduated cylinder—25, 50, or 100—ml
large container of water; cloth bag; spoon
string about 6–10 inches long
egg yolk; fat or oil; flour—all purpose or bread
sucrose (cane sugar)

What to do

PART I THE PROTEINS IN EGG WHITES

1. Pour 100 ml–150 ml of water into a 250–ml beaker.
2. Assemble an iron stand with ring, a wire gauze, and a burner. Set the beaker on the wire gauze and heat the water until it boils.
3. Turn off the burner, and drop the white of an egg into the hot water. Observe the egg white for two–four minutes. What do you observe?

   ____

   ____

4. Remove a bit of the egg white with forceps. Place it on a watch glass. Set the watch glass aside for use in step 6.
5. Relight the burner and boil the remaining egg white for two–three minutes.
6. Remove a bit of the egg white and compare it with the sample you set aside in step 4. What do you observe? ______________________________

   ____

   ____

What evidence indicates that the proteins in the egg white have coagulated? ____________________

____

____

____

PART II THE PROTEINS IN MILK

1. Pour about 10 ml of skim milk into a test tube.
2. Add dilute hydrochloric acid with a dropper until a noticeable change occurs. What happens?

   ____

3. Let the contents of the test tube stand for a few minutes. Carefully decant the liquid.

   Describe the appearance of the residue? ________

   ____

   (a) What dairy food does the residue resemble?

   ____

   ____

   (b) During the souring of milk, lactic acid is formed. Why does sour milk curdle? ________

   ____

PART III THE PROTEINS IN FLOUR

1. Place about 50 grams of flour in a large evaporating dish or saucer.
2. Add 25 ml of cold water, and stir the flour and water mixture until all of the flour has become moist.
3. Dust flour on your hands, then transfer the dough to your hand.
4. Knead the dough until it becomes elastic and smooth.

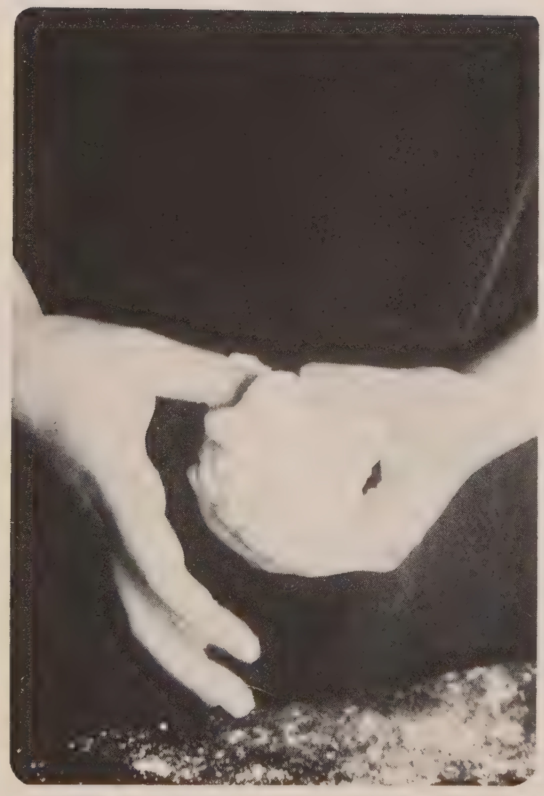


Fig. 88. *(Clara Aich and George Senty)*

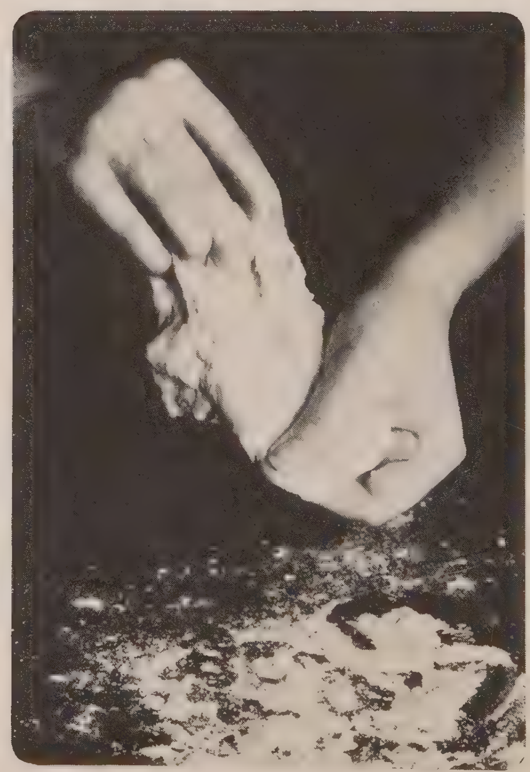


Fig. 89. *(Clara Aich and George Senty)*

Fig. 90. *(Clara Aich and George Senty)*

The kneading takes from three to five minutes.

5. Place the ball of dough in a cloth bag and tie the bag with a string.
6. Immerse the bag in a container of cold water. Work the water through the dough to remove the starch. You will need to change the water when it gets cloudy and white.
7. When the dough no longer makes the clear water cloudy, place the bag under running cold water for one or two minutes.
8. Open the bag and inspect the residue of coagulated proteins. This residue is called GLUTEN. Describe the appearance of the gluten. ______________________________

   ____
9. Scrape the residue from the bag. Form the residue into a wad of gluten. Gently try to pull the wad apart. Record any further observations you make. ______________________________

   ____

   ____

   ____

Extra for experts

1. Do sugar, fat or oil, and egg yolks retard or increase the formation of gluten in doughs? Plan an experiment to find the answer to the question. After the teacher has approved your plan, have some of your classmates help you carry out the experiment. Briefly describe your plan. ______________________________

   ____

   ____

   ____

   What did you find? ______________________________

   ____

   ____
2. Read the directions and ingredients listed in bread and cake recipes. Account for the difference in the tender texture of cakes and firm texture of bread and cream puffs.

KITCHEN COLLOIDS

An Experiment 16.3, ABOUT STARCH, you learned that large clumps of molecules may be spread out evenly in a solvent. The clumps are large enough to scatter light and so show the pathway of a beam of light. If large particles remain spread out evenly in another material, and do not settle out even after a long time, they are said to be in the COLLOIDAL STATE. They form a COLLOIDAL DISPERSION. A true solution is clear. A colloidal dispersion may be cloudy.

There are four important kinds of colloids: SOLS, EMULSIONS, GELS, and AEROSOLS. To these might be added foam and sponge.

Sol ------------ solid suspended in a liquid (milk of magnesia)
Emulsion --- liquid suspended in a liquid (homogenized milk)
Gel ----------- liquid suspended in a solid (jello)
Aerosol
Smoke ------- solid suspended in a gas (smoky air)
Fog ----------- liquid suspended in a gas (foggy air)
Foam -------- gas suspended in a liquid (whipped cream)
Sponge ------ gas suspended in a solid (plastic foam)

Many of the substances found in the kitchen are colloidal dispersions. See how many of the colloids assigned to you, you can identify.

Experiment 17.4

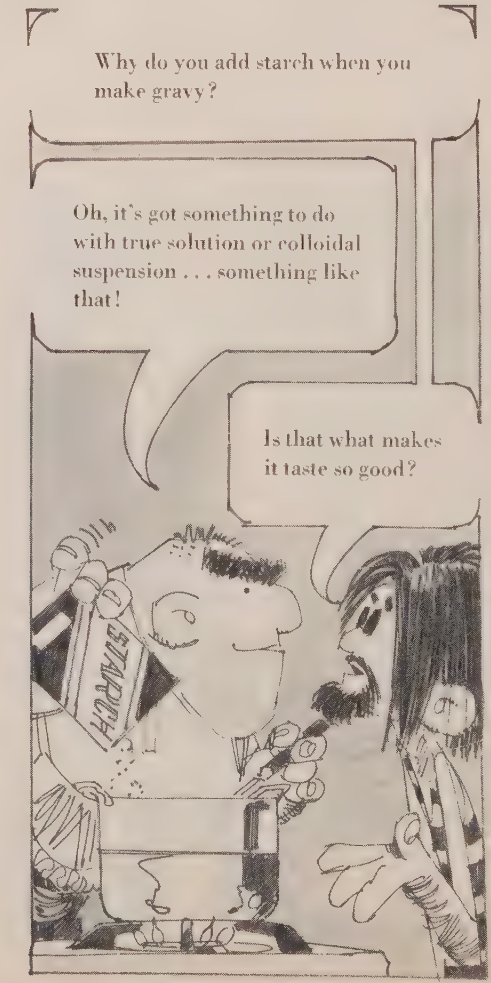
Equipment and Supplies per Lab Station
bowl—quart size or liter beaker
graduated cylinder; metal tablespoon
cornstarch; samples of colloids

What to do

PART I

1. Inspect each of the colloids and discuss them with your partner. Classify them according to the above groups. (Hint: Sometimes finding out how the substance is made will help you discover the kind of colloid.)

2. As you classify each colloid, complete the information in the Data Chart.


Why do you add starch when you make gravy?
Oh, it's got something to do with true solution or colloidal suspension . . . something like that!
Is that what makes it taste so good?
STARCH

DATA CHART

Name of mixture	Name of colloid	Suspended material	Material doing the suspending
1.			
2.			
3.			
4.			
5.			

PART II A TRICK TO FOOL YOUR FRIENDS

1. Get a bowl or beaker large enough so that you can put your hand into it to remove something.
2. Put a cup of cornstarch, about 250 ml, into the bowl. What is the appearance of the cornstarch? ________________
3. Pour 10 ml of water into the bowl. Stir the mixture with a spoon. Describe what happens. ________________ ________________
4. Continue to add water until the cornstarch and water begin to shine. Describe what happens as you stir the mixture. ________________ ________________
5. Continue to add water until the stirring is difficult. The mixture should seem to be very wet and fluid when you stop stirring.
6. KEEP YOUR HAND OVER THE SINK ALL THE TIME YOU ARE DOING THIS STEP. Put some of the mixture in the beaker into your hand. IMMEDIATELY press the material. Use both hands. Try to roll it into a ball. Does a ball seem to be formed? ________________

   Stop rolling and pressing the starch. Hold your hand flat. What happens now? ________________ ________________
7. Again make a ball by constantly rolling the starch around in your hand. Ask some unsuspecting person to hold the cornstarch ball.

   Watch his face as the ball begins to disappear!

A WORD OF ADVICE—If you try this at home, keep your hands and the cornstarch ball over the sink. It's pretty messy to clean up!

Questions to answer

1. Why does starch and water show the properties you have just observed? ________________ ________________ ________________
2. Can you think of something outdoors that shows the same properties? ________________ ________________

TAKING THE **PLOP** OUT OF KETCHUP

Have you ever noticed how difficult it seems to get some ketchups out of the bottle? You shake and shake and then suddenly, PLOP! Ketchup is all over the place.

Ketchup is a colloid in which the suspended solids trap the liquid medium. When disturbed by vigorous shaking the solid particles break away from the arrangement which trapped the liquid. As a result of the shaking, the colloid becomes very fluid. The colloid may return upon standing to the way it was originally.

Demonstration 17-A

Equipment and Supplies per Lab Station

beaker—100 ml; blender
ketchup, grocery store bottle of ketchup

What to do

1. Uncap a fresh bottle of ketchup. Turn it upside down. Does any ketchup come out of the bottle?

   ____

2. Put the cap on. Shake the bottle vigorously. Take off the cap. Then hold the ketchup over a blender container. Does any ketchup come out now?

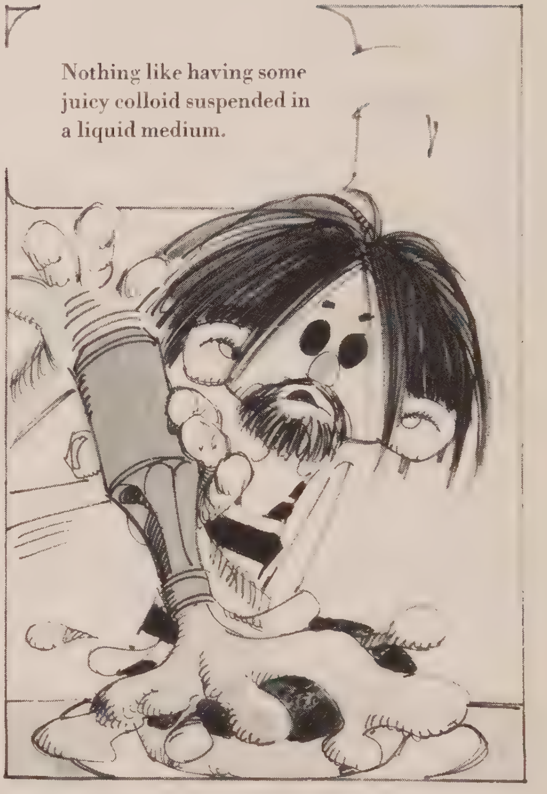
   ____

3. Continue to shake the bottle. Transfer all of the ketchup to the blender container. Put the blender on low speed for 10 seconds. Then put one spoonful into a 50 or 100-ml beaker. Set the beaker aside. Label the beaker: 1.
4. Turn the blender on high speed for 10 seconds again. Remove another spoonful to a beaker marked: 2. Set the beaker aside.
5. Turn the blender on for three minutes at high speed. Remove another spoonful to a beaker marked: 3.
6. Compare the appearance of the three samples of ketchup taken from the blender. Describe their appearance. Test the way they can be poured from the beaker.

   No. 1 ______________________________

   No. 2 ______________________________

   No. 3 ______________________________


Nothing like having some juicy colloid suspended in a liquid medium.

Questions to answer

1. Why can you remove ketchup from a bottle which is only 1/4-full easier than from a full bottle? ______________________________
2. What is the name for a colloid in which a solid is suspended in a liquid? ______________ Look up directions for making ketchup. What do you think the suspending liquid is in ketchup? _____

   ____

3. Can you think of any other substance in the kitchen which reacts the same way to beating or stirring that ketchup does? ______________ Test the material by stirring it with a spoon and noting any change in its appearance. _____

   ____

A COLORFUL SEPARATION

Natural food coloring and other substances are often added to food. They are called FOOD ADDITIVES. These food additives improve the appearance or taste, or they retard spoilage of the food.

Sometimes foreign matter such as pesticides, fertilizers, and other harmful substances appear in our food. They are called FOOD CONTAMINATORS.

There are chemists who are on the lookout for additives or contaminators which can be harmful to man. How can the presence of these undesirable substances be detected? One technique often used is called CHROMATOGRAPHY.

CHROMATOGRAPHY is a method used to separate parts of a mixture, so the component parts can be identified by their colors. A spot of the mixture is placed on a substance called an ADSORBENT. Paper and a thin layer of silica gel are examples of adsorbents. A solvent moves along the adsorbent and takes some or all of the components away from the original spot. The component parts of the mixture separate since they do not all move along the adsorbent at the same rate.

How many components are in the food coloring mixture given you? Make a chromatogram and find out. *A CHROMATOGRAM is the pattern in in the adsorbent (paper or silica gel) which results when the material to be separated is dissolved in a solvent and allowed to move along the adsorbent.*

Equipment and Supplies per Lab Station

capillary tube; developing jar with stopper; metric ruler; pencil
(3) strips Whatman #1 filter paper
(1) strip Whatman # 3 filter paper or silica gel thin layer
ethanol; ethanol-water mixture 1:1; unknown mixture of food colorings
water

Experiment 17.5

What to do

PART I

1. Obtain the following equipment and supplies: a developing jar, a strip of Whatman #1 filter paper, capillary tube, and an unknown mixture of food coloring.
2. Draw a pencil line about 2 cm above one end of the paper strip.
3. Dip the capillary tube into the dye mixture. After some dye has moved up into the tube, quickly touch the capillary tube to the line on the paper strip. A small spot of dye will appear on the paper. Let the spot of food coloring dry.
4. Pour water to a depth of 1 cm in the developing jar.
5. Fasten the strip of paper to the cover of the jar as shown in Fig. 91.
6. Insert the strip of paper into the jar so the end nearest the spot is in the water. The spot must not be immersed!
7. Watch the chromatogram develop as the water begins to move up the paper. Write down all that you observe.

____

____

____

____

8. After the time specified by your teacher, remove the chromatogram. How many separate colored areas do you see? ______
How many food dyes do you think are present in the mixture? ______
Where do you think you would find a component that could not dissolve in the solvent? ______
______

PART II SOME LABORATORY PROBLEMS TO SOLVE

1. Does the kind of adsorbent used affect the separation of a mixture? Repeat steps 1–8 in PART I using Whatman #3 filter paper or a strip of thin layer silica gel for the adsorbent.
   (a) Write down all of your observations as the chromatogram develops.
   ______
   ______
   ______
   ______
   ______
   ______
   (b) What is your answer to the above problem?
   ______
   ______
   ______
   ______

2. Does the kind of solvent used make a difference in the separation of the components?
   ______ Repeat steps 1–8 of PART I but substitute ethanol for the water as a solvent.
   (a) What do you observe as the ethanol moves along the paper?
   ______
   ______
   ______
   ______
   (b) What is your answer to the above problem?
   ______
   ______
   ______
   ______

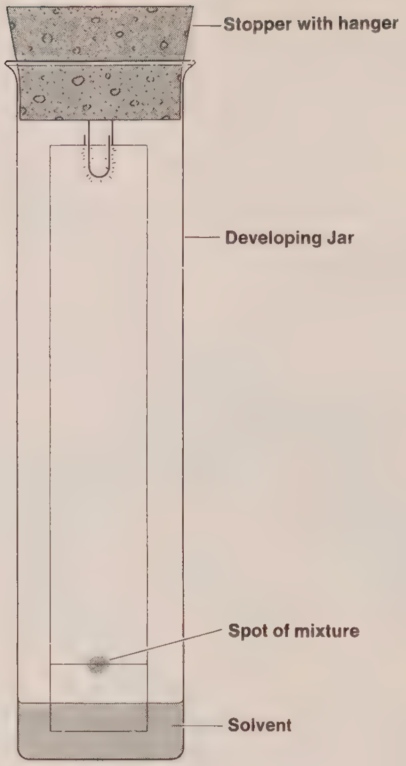


Fig. 91.

Extra for experts

Do you think that you would get a separation of the components if you used a solvent mixture containing equal parts of water and ethanol?
______

Experimentally prove or disprove your answer by making the chromatogram using the Whatman # 1 adsorbent and the solvent mixture.
______
______
______
______
______

What did you find out? ______
______
______

Calculating and Using R_f Values

Exercise

In Experiment 17.5 you found that the components of a mixture traveled at different rates along an adsorbent. Both the kind of adsorbent and kind of solvent used had an effect on the rate at which each component flowed. Chemists make use of these differences in rates to identify components.

Identification is based upon rate of flow or R_f values. A chromatogram is made using an appropriate adsorbent and solvent. Then the values are calculated from the formula:

$$R_f = \frac{\text{distance the component moved}}{\text{distance the solvent moved}}$$

Study the chromatogram below, then answer the questions.

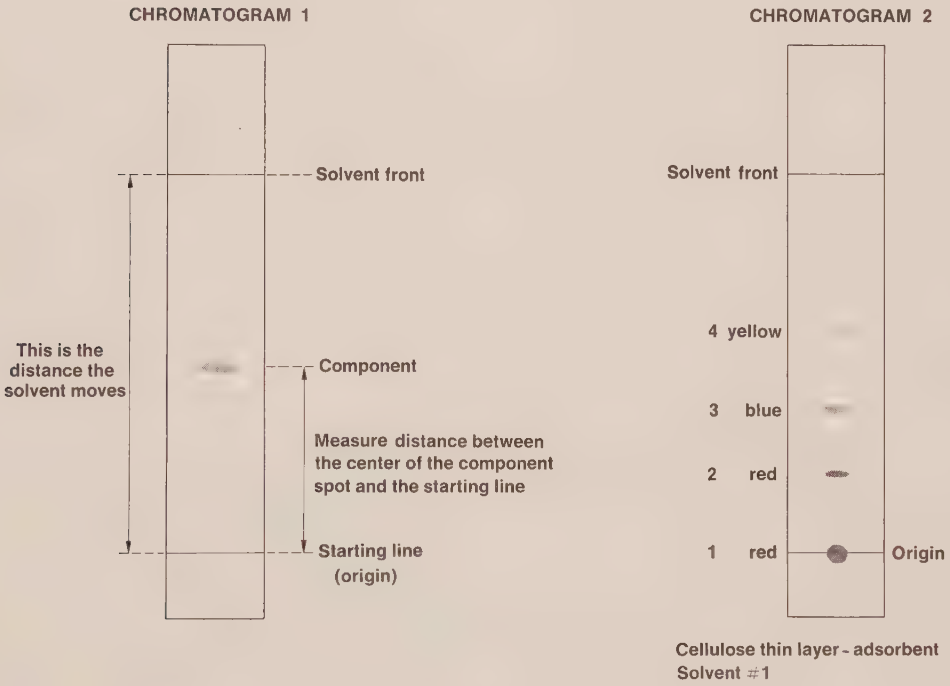

CHROMATOGRAM 1
Solvent front
This is the distance the solvent moves
Component
Measure distance between the center of the component spot and the starting line
Starting line (origin)
CHROMATOGRAM 2
Solvent front
4 yellow
3 blue
2 red
1 red
Origin
Cellulose thin layer - adsorbent
Solvent #1

Fig. 92.

Use your ruler to measure the two distances:

The component moved ________________ cm

The solvent moved ________________ cm

Calculate the R_f value using these values:

$$R_f = \frac{\text{distance the component moved}}{\text{distance the solvent moved}}$$

$$R_f = \frac{\qquad\qquad \text{cm}}{\qquad\qquad \text{cm}},$$

$R_f =$ __________ answer

The R_f value you just calculated shows that the component moved half as fast as the solvent did.

Now calculate the R_f values for each of the components on Chromatogram 2. Write your answers in the Data Chart below.

DATA CHART

Component No.	Component distance	Solvent distance	R_f value
1			
2			
3			
4			

In the APPENDIX you will find an R_f Reference table. By matching up colors and R_f values, identify as many of the components as you can.

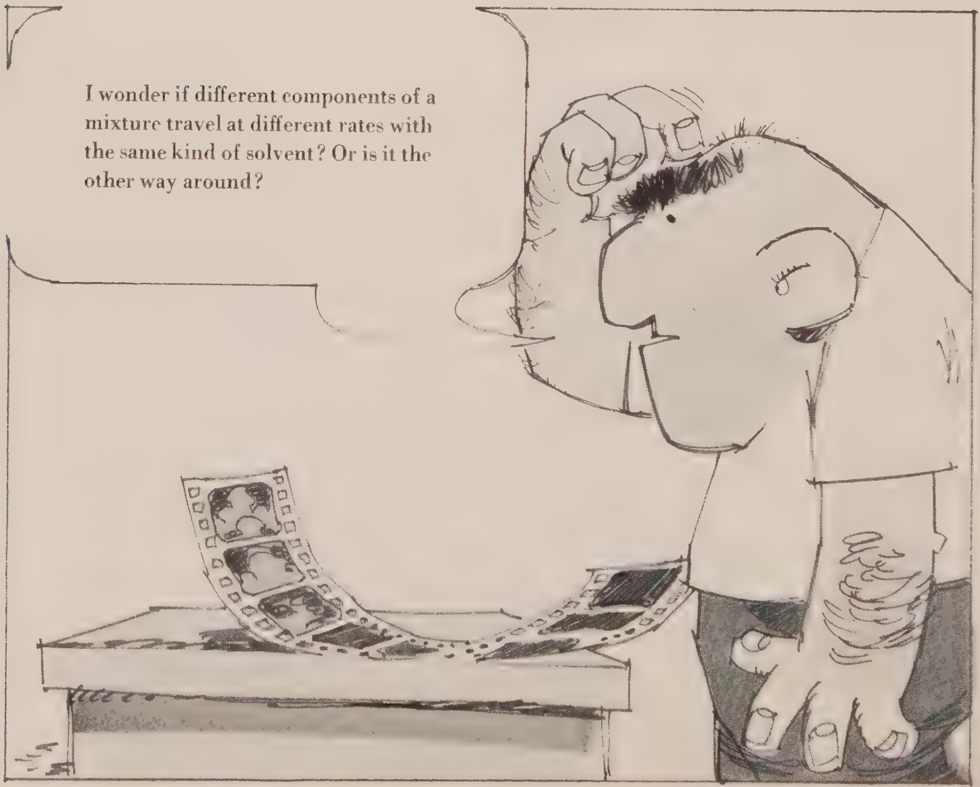
What components were shown to be present on the chromatogram? ____________________

____________________ ____________________

Could you use this Reference Table if you used a different solvent? ______________ Why? ________

____

______________ ____________________________


I wonder if different components of a mixture travel at different rates with the same kind of solvent? Or is it the other way around?

Making Soap

Soap is something you take for granted; but what is it made of? Make some and find out!

Equipment and Supplies per Lab Station

beaker—150 ml; beaker—250 ml; evaporating dish
graduated cylinder
burner; ring stand with ring and wire gauze
spatula; stirring rod, glass
cottonseed oil, 10 ml; ethanol, 10 ml
sodium chloride, saturated solution, 25 ml
sodium hydroxide, 30%, 5 ml
soap, uncolored, small piece from a bar of soap

What to do

1. Pour 10 ml of cottonseed oil into a clean evaporating dish.
   Note: If your evaporating dish is too small to hold 100 ml of solution use a 150—ml beaker.
2. Add to the oil, 10 ml of ethanol and 5 ml of 30% sodium hydroxide solution.
3. Heat the mixture *gently, with constant stirring,* until the material becomes thick.
4. Continue heating *gently* until you can see no more droplets of oil on the surface of the mixture. Watch carefully. The mixture froths over easily. This heating should take 10 to 12 minutes.
5. Set the warm evaporating dish in the top of a 250—ml beaker filled with *cold* water.
6. Add 20 ml of *hot* distilled water. Stir well to mix.
7. The water in the beaker is now warm. Replace it with cold water. Return the evaporating dish to the top of the beaker.
8. Add 25 ml of a saturated solution of sodium chloride. Stir to mix. Let stand for about two minutes.
9. Scoop off the curds of soap onto a paper towel. Form the soap into a small bar. Wash the liquid left in the evaporating dish down the sink.
10. Compare your soap with a piece from a bar of uncolored soap.
11. Save your soap for Experiment 18.2.

	Color	Hardness	Soapy feeling	Suds on damp fingers
Your soap				
Soap bar				

Questions to answer

1. Name three or more substances which might have been used instead of the cottonseed oil.

____

____

____

2. What substance might have been used instead of the sodium hydroxide? ______________
3. Why was the sodium chloride solution added?

____

____

4. How did your soap compare with the commercial soap? ______________

____

MAKING A SYNTHETIC DETERGENT

Most of the "washday products" are synthetic detergents. There are many different brands, all highly advertized.

Watch for differences from soap making in the materials and method of preparation in this demonstration.

Demonstration 18-A

Equipment and Supplies per Lab Station

beaker—150 ml; beaker—250 ml
Büchner funnel with filter paper to fit
filter flask, heavy walled to withstand pressure
graduated cylinder; medicine dropper; stirring rod, glass
aspirator (filter pump)
lauryl alcohol (dodecanol), 18.6 g; sodium hydroxide, 2*M*, 35 ml
sulfuric acid, concentrated, 7 ml; phenolphthalein
in a dropper bottle

What to do

1. Place 18.6 g of lauryl alcohol (dodecanol) in a 150-ml beaker. Add very *slowly* and *carefully* 7 ml of concentrated sulfuric acid.
2. Transfer 35 ml of 2*M* sodium hydroxide solution to a 250-ml beaker. Add three to four drops of phenolphthalein. (The phenolphthalein may not turn red in this concentrated sodium hydroxide solution.)
3. *Cautiously* pour the acid solution from Step 1 *slowly* and *with constant stirring* into the sodium hydroxide solution. The phenolphthalein will change from colorless to red to colorless. If it does not become colorless, add sulfuric acid drop by drop until it does. A thick white precipitate forms.
4. Filter through a Büchner funnel, using an aspirator as illustrated. Allow the product to dry.

   The product may be used in Experiment 18.2 where the properties of soaps and synthetic detergents are compared.

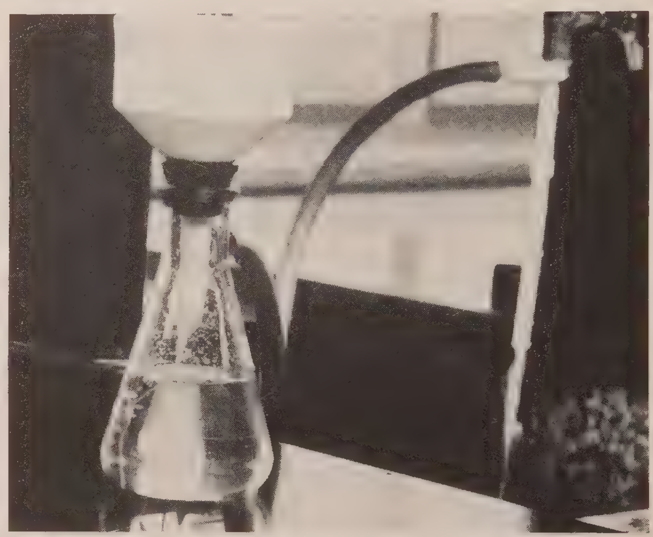


Fig. 93. *(Clara Aich and George Senty)*

Questions to answer

1. What starting substance was used instead of cottonseed oil? ______________________

   ____
2. What compound was used both in making soap and in making this synthetic detergent? ______

   ____
3. (a) How was the soap separated from the solution in which it was made? ______________

   ____

   (b) How was the synthetic detergent separated from the solution in which it formed? ______________

   ____
4. Compare a small amount of the dry synthetic detergent with the soap you made.

   (a) Likenesses: ______________

   ____

   ____

   (b) Differences: ______________

   ____

   ____

WETTER WATER

Water by itself is not wet enough for many cleaning jobs. Detergents, both soaps and the synthetic types, make water wetter and then they power in and get the dirt out.

What do you mean, "wetter water?" What is SURFACE TENSION?

SURFACE TENSION: The molecules in water like to stick close together. They form sort of a skin on the surface of the water. Rain and poured water form drops. Water on a glass or metal surface forms large drops instead of spreading out evenly. *The pull molecules on the inside have on molecules toward the surface is called SURFACE TENSION.*

Detergents and the surface tension of water: Detergents break up the drops so that water spreads out more evenly. That is, detergents "reduce surface tension." This makes water wetter.

Experiment 18.2

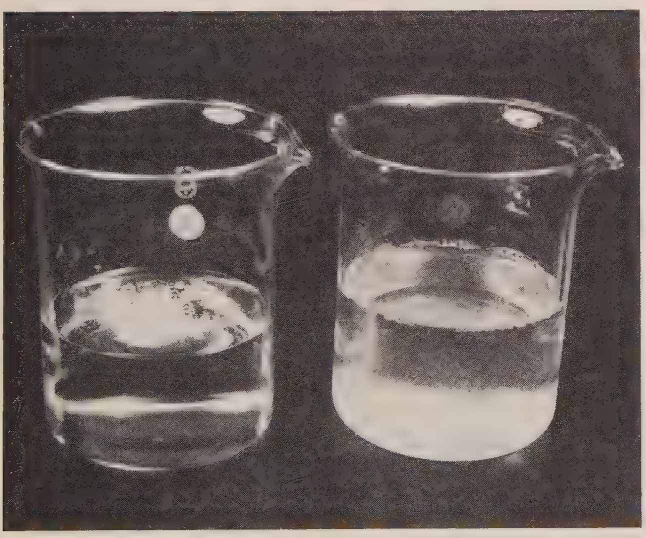


Fig. 94. *(Clara Aich and George Senty)*

Equipment and Supplies per Lab Station

(3) beakers—150 ml; (2) glass plates
graduated cylinder, 25 ml or 50 ml; metric rule
(4) test tubes; Universal Indicator paper
(4) small balls of nonabsorbent cotton; distilled water
hydrochloric acid, concentrated, 3 ml; oil, cottonseed, 1 ml
oil, mineral, 1 ml; small sample of pure soap
small sample of synthetic detergent

What to do

PART I

1. Place a tiny bit of the soap you made in a test tube. Half fill the test tube with *distilled* water. Shake vigorously for 15 to 20 seconds. Examine the test and record the following in Data Chart 1:
   (a) the height of the suds layer measured in centimeters.
   (b) the amount of the soap that seemed to dissolve.
   (c) whether the solution is transparent or not.
2. Repeat step 1, using *tap* water instead of distilled water.
3. Repeat steps 1 and 2, using the synthetic detergent prepared in Demonstration 18-A or a commercial product.

DATA CHART 1

	Distilled Water			Tap Water		
Sample	Height of suds (cm)	Amount dissolved	Transparent or not	Height of suds (cm)	Amount dissolved	Transparent or not
Soap						
Synthetic Detergent						

PART II

1. Dissolve, without shaking, a small amount of your soap in a 150–ml beaker half-filled with distilled water. Determine the properties of soap in steps 2 through 6. Record in Data Chart 2.

2. *pH*: Determine the pH, using Universal Indicator paper.

3. *Surface tension*: Dip a small ball of cotton in mineral oil. Use this to spread the oil on a glass plate so as to make a thin film.

   (a) Place a drop or two of distilled water on the plate. Record the measurement, in centimeters, of the approximate diameter of the drop.

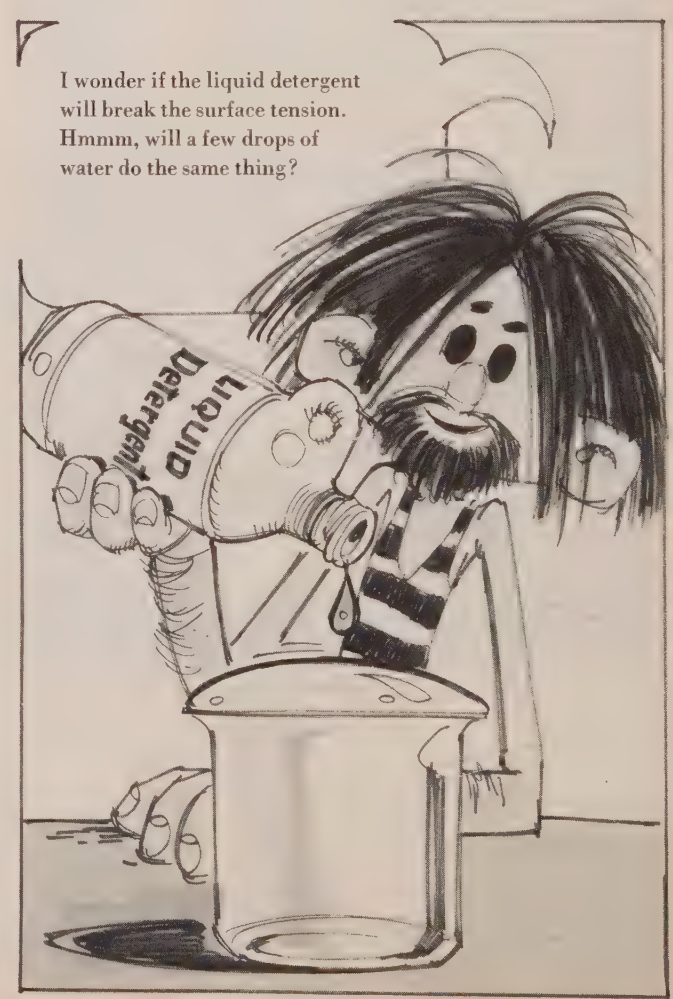
   (b) Repeat, using the same number of drops of your soap solution. Record high, moderate, or low surface tension.
   High surface tension – forms drop which can be measured.
   Low surface tension – spreads out.

4. *Wetting ability*
   (a) Half fill a beaker like the one containing your soap solution, with *distilled* water. Drop a small ball of nonabsorbent cotton on the surface of the water. Record what happens.

   (b) Drop another small ball of nonabsorbent cotton on the surface of your soap solution. Again record what happens.

5. *Emulsifying effect*
   (a) Place one or two drops of cottonseed oil in a test tube. Half fill the test tube with distilled water. Shake vigorously. Let stand two minutes. Record the result.

   (b) Repeat (a) using your soap solution instead of the distilled water. Record the result.


I wonder if the liquid detergent will break the surface tension. Hmmm, will a few drops of water do the same thing?
Liquid Detergent

6. *Effect of hydrochloric acid, HCl, on soap solution:* Remove the cotton ball from the remainder of your solution in the beaker. Add 3 ml of concentrated hydrochloric acid. Record what happens.

7. Dissolve, without shaking, a small amount of synthetic detergent in a 150 – ml beaker half-filled with *distilled* water. Repeat steps 2 through 6 using this solution. Record the results in Data Chart 2

DATA CHART 2

Solution	pH	Surface tension	Wetting ability	Emulsifying effect	Effect of HCl
Distilled Water					
Soap Solution					
Synthetic Detergent					

Questions to answer

1. What differences, if any, did you observe between soap and synthetic detergent in
   (a) amount of suds ____________
   (b) amount dissolved ____________
   (c) transparency of the solution ____________
   ____

2. Excess base in a washing product can damage delicate fabrics. Would you recommend soap or synthetic detergent to clean these fabrics? Why?
   ____
   ____

3. How do you explain any difference you may have observed in the way the cotton ball acted in distilled water, soap solution, and synthetic detergent solutions? ____________
   ____
   ____

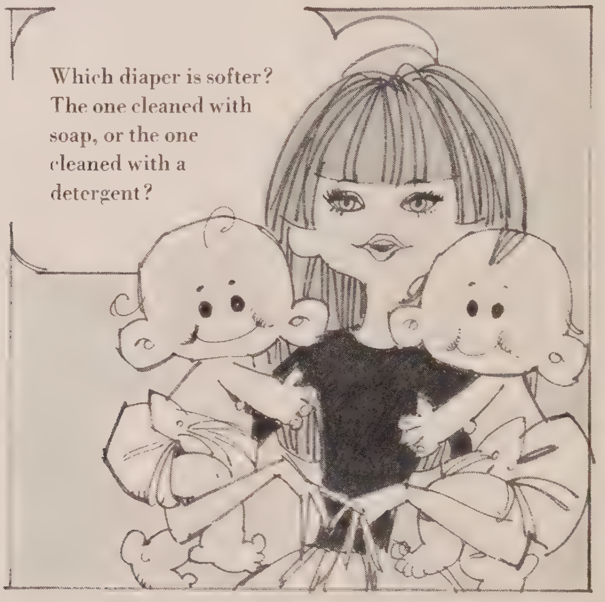
4. In Experiment 16.4 you tested the effect of egg albumin on a mixture of oil and water. Does soap or synthetic detergent have a similar effect on fat? Explain. ____________
   ____
   ____
   (a) How does the effect on fat help the cleaning action of soap or of synthetic detergent? ____________
   ____
   (b) Which works better, the soap or the synthetic detergent, if there is a difference?
   ____

5. What is the effect of hydrochloric acid on
   (a) soap solution ____________
   (b) synthetic detergent solution ____________
   (c) How might this reaction affect their cleaning properties? ____________
   ____
   ____

Soap or Synthetic Detergent?

Use your experience in Experiment 18.2 to test unknown samples. Decide whether each one tested is a soap or a synthetic detergent.

Experiment 18.2-A Extension


Which diaper is softer? The one cleaned with soap, or the one cleaned with a detergent?

Equipment and Supplies per Lab Station

Equipment and supplies as in Experiment 18.2
samples of unknowns

What to do

1. Select at least three unknown samples.
2 Make any tests you think are appropriate and necessary. Record the results. Decide whether the sample is soap or synthetic detergent. Give your reasons.

UNKNOWN SAMPLES

Sample	Tests Made	Results	Soap or Synthetic	Reason
#1				
#2				
#3				
#4				

3. Was your decision wrong in any case? ________

If so, why? ____________________

____

____

Effect of Disinfectants on Bacterial Growth

Many cleaning agents are advertised for their ability to destroy household germs. How effective are they?

Demonstration 18-B

Equipment and Supplies per Lab Station

(6 to 8) nutrient agar culture plates, sterile

(4) different cleaning agents advertised for their ability to destroy household germs

What to do

1. Obtain six to eight sterile, nutrient agar culture plates. Label each one by placing a number on a sheet of paper under each one.
2. *Control*: Keep the first plate unexposed.
3. Remove the cover from each of the other plates and leave the agar exposed for two minutes.
4. Add two drops of tap water to one plate. Replace the cover.
5. Add two drops of a different disinfectant solution to each of the other plates. Replace the covers. Let the cultures set for two to three days. Be sure to keep them covered. Place them where they will not be disturbed.
6. Observe the bacterial growth in each culture every day.
   Record what you see.

RECORD OF GROWTH

Disinfectant	24 hours	48 hours	72 hours
1. Control			
2. Tap water			
3.			
4.			
5.			
6.			

Questions to answer

1. Was there any growth in the control? ____________
2. Is tap water a good disinfectant? Explain. ________

____

____

3. (a) Which of the disinfectants tested seemed to be the most effective? ________________
   (b) Which seemed to be the least effective?

____

____

Chapter 19

THE SOFTEST MINERAL

Talc is the softest mineral known. It has also been thought to be harmless to the skin. For these reasons talc is used as the base of most cosmetic powders.

Talcum powder is sifted powdered talc. Perfume is added to commercial talcum powder.

Face powders are made from a combination of powdered minerals which do not harm the skin. Talc is one of the minerals. Commercial face powders are perfumed and colored to match skin tones.

Experiment 19.1

Equipment and Supplies per Lab Station

(2) beakers—250 ml; glass plate; mortar and pestle
glass stirring rod; test tube; balance
hand lens or microscope
material, silk or other lint-free type, 3 in × 3 in
calcite, $CaCO_3$; limestone, $CaCO_3$
0.6 g precipitated chalk, $CaCO_3$; 1.6 g kaolin
3.0 g powdered talc; 1.0 zinc oxide, ZnO
0.6 zinc stearate, $Zn(C_{18}H_{35}O_2)_2$; hydrochloric acid, 1*M*, HCl

The softest mineral on earth. Soft enough for a baby's skin. Is it soft enough for your baby's skin?

What to do

PART I TALC

1. Examine a piece of talc rock.
   (a) What color is it? ________
   (b) Rub it between a thumb and forefinger. How does it feel? ________
   (c) Does some of the talc stick to your skin? ________
   (d) Will the talc make a mark on cloth or other substances? ________
   (e) Can you scratch the talc rock with your thumbnail? ________
2. Grind a small amount of talc rock in your mortar. Look at the particles under a hand lens or a microscope. Describe what you see. ________
3. Place a small amount of the powdered talc in a test tube. Add a few drops of 1*M* hydrochloric acid. Is there any bubbling? ________
   Note: If there is bubbling, carbonates may be present. Carbonates are an impurity in talc rock.
4. Compare the hardness of talc rock with that of calcite and limestone. Scratch one on the other. The one scratched is always the softer one. Arrange the three minerals in the order of increasing hardness.
   (a) Softest ________
   (b) Next softest ________
   (c) Hardest ________

5. Place a piece of each of the minerals on a glass plate. Drop on one or two drops of 1*M* hydrochloric acid. Check the ones that bubble.

   (a) Talc ________ (b) Calcite ________ (c) Limestone ________

PART II TALCUM POWDER

1. Sift a small amount of powdered talc rock through silk as illustrated.
2. Rub a little of the sifted talc on the skin of your wrist.
   (a) How soft does it feel? ________
   (b) Does it cling smoothly to the skin? ________
3. Wash all powdered talc off your skin.

PART III FACE POWDER

1. Place in a mortar:
   3.0 g powdered talc
   1.0 g zinc oxide
   0.6 g zinc stearate
   0.6 g precipitated chalk
   1.6 g kaolin
2. Grind the minerals together. Mix thoroughly. Sift the face powder as you did the talcum powder.

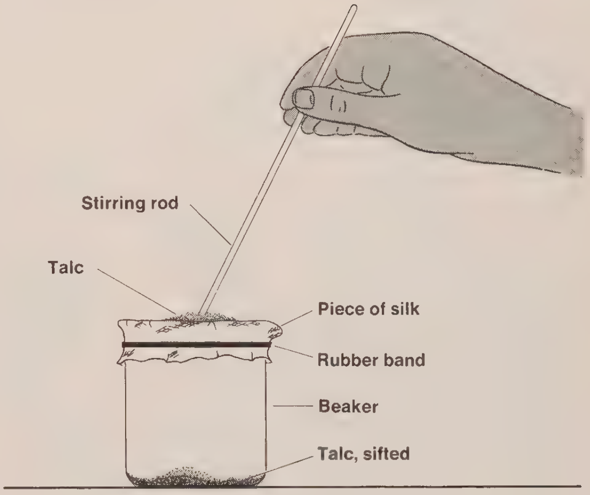


Fig. 95.

Questions to answer

1. How is face powder different from talcum powder? ________

   ____
2. What does bubbling indicate when hydrochloric acid is dropped on calcite and limestone? ________

   ____
3. Limestone is calcium carbonate. It reacts with hydrochloric acid. Write the symbols and balance the equation.

   calcium carbonate + hydrochloric acid → calcium chloride + carbon dioxide + water
4. Look up talc in a reference book.
   (a) What chemical compound is it? ________
   (b) How is it obtained? ________
   (c) Where is it found in the United States?

   ____

Yes sir, I use calcium carbonate to keep my teeth bright and clean. Yes sir, wouldn't use anything else, but good old calcium carbonate.
YES SIR, I USE CALCIUM CARBONATE TO KEEP MY TEETH BRIGHT AND CLEAN YES SIR
TAKE 92

TOOTH POWDER AND TOOTHPASTE

Tooth powder or toothpaste contains a mild scouring agent which will not damage enamel, and a substance which will leave the mouth slightly basic in reaction.

Tooth powders usually are made from precipitated chalk, which is calcium carbonate. Flavored precipitated chalk can be purchased at drugstores.

Toothpastes contain precipitated chalk, but other ingredients are added. One of these may be a small amount of soap which can foam away food particles. The dry materials are moistened to make a soft paste. Toothpastes are generally flavored slightly.

Experiment 19.2

Equipment and Supplies per Lab Station
beaker; glass bend, 60° angle; evaporating dish
flask—125 ml; flask—250 ml
thistle tube or long-stemmed funnel; funnel with filter paper
mortar and pestle; stirring rod, glass
glass tubing, 3 inches; soda straw or 6 inches of fire-polished glass tubing
filter rack or ring stand with small ring
rubber stopper, 2 hole; rubber or plastic tubing, 3 inches
red and blue litmus paper; 1.5 g powdered white soap
5 g marble chips; 7.0 g precipitated chalk
5.0 g magnesium carbonate; 1.5 g powdered sugar
limewater, 150 ml; glycerin solution, 2—5 ml
dilute hydrochloric acid, 100 ml

What to do

PART I TOOTH POWDER

1. Make precipitated chalk
   (a) Set up a carbon dioxide generator. See Fig. 96. Bubble the carbon dioxide released through 50 ml of limewater until the limewater becomes very milky.
   Note: Your teacher may direct you to bubble your breath through the 50 ml of limewater until the limewater becomes very milky instead of using the carbon dioxide generator.
   (b) Filter out the precipitated chalk. Remove the filter paper from the funnel. Spread it out in an evaporating dish to dry the precipitate.
   Note: Start PART II while the precipitate is drying.
2. Examine the dry precipitated chalk toward the end of the class period. Rub a little between a thumb and forefinger.

   (a) Is it smooth or gritty? ______________________

   (b) How hard does it seem? ______________________

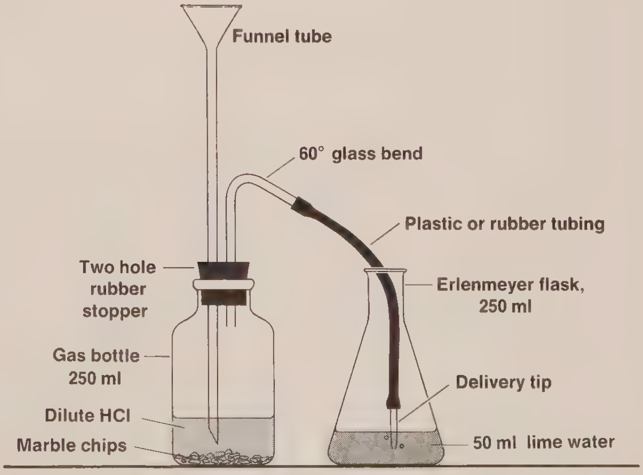


Fig. 96.

3. Place some of the precipitated chalk in an evaporating dish. Moisten it with a few drops of water. Try using it to scour the dish. How does it work? ________________

CAUTION: DO NOT TRY ANY ON YOUR TEETH! LAB CHEMICALS MAY NOT BE PURE.

PART II TOOTHPASTE

1. Grind together in a mortar:
   7.0 g precipitated chalk
   5.0 g magnesium carbonate
   1.5 g powdered sugar
   1.5 g powdered white soap
2. Add flavoring if your teacher directs.
3. Add glycerin solution until a soft paste is formed. About 1 to 3 ml should be enough.
4. Place a little of your toothpaste in an evaporating dish. Add a small amount of water. Mix well.
   (a) Is the solution acid or basic to litmus?
   ____
   (b) Use a test tube brush and the toothpaste to scrub the evaporating dish. Rinse and blot dry. Can you find any scratches on the surface of the dish? ________________

CAUTION: DO NOT TRY ANY ON YOUR TEETH! LAB CHEMICALS MAY NOT BE PURE.

Questions to answer

1. Write the symbols and balance the equation for the reaction between limewater (calcium hydroxide solution) and carbon dioxide.

calcium hydroxide + carbon dioxide → calcium carbonate + water

2. Many dentists have recommended the use of a mixture of one part by weight of powdered table salt and two parts by weight of baking soda as a tooth powder.
   (a) Which of the two components would give the scouring action? ________________
   (b) Which would tend to leave the mouth slightly basic in reactions? ________________
3. Would you expect toothpaste to neutralize acid residues from food in the mouth? Explain. ________________
   ____

Irritating Substances

Experiment 19.2-A Extension

Equipment and Supplies per Lab Station

glass stirring rod; (8) test tubes; test tube rack
2 to 5 strips of litmus paper, red and blue
2 to 5 strips of tumeric test paper
1 ml of borax solution
10 ml of potassium chlorate solution, $KClO_3$
7 ml of conc. hydrochloric acid, HCl
samples of 3 different toothpastes (include some with whiteners)

Toothpaste may contain borax, perborate, or chlorate. These substances can be irritating to gums around the teeth, especially after continued use.

Test for the presence of boron compounds and chlorates in a number of samples of commercial toothpaste. Record your results in the Data Chart.

What to do

1. *Test for boron compounds*
   - (a) *Control*: Moisten a strip of tumeric paper in a solution of borax. Dry the paper in a warm place.

     What color develops in the tumeric paper?

     ____

     Note: This color change in tumeric paper shows the presence of a boron compound.
   - (b) Place a small amount of the toothpaste to be tested in a test tube. Add 2 to 3 ml of water. Shake well. Moisten a strip of tumeric paper in the solution. Again, dry the paper in a warm place. Record the color change, if any, in the Data Chart.
2. *Test for chlorates*
   - (a) *Control*: Half fill a test tube with a solution of potassium chlorate. Add 7 ml of concentrated hydrochloric acid. What color appears? ____________________

     Dip a strip each of red and blue litmus paper in the solution. What happens to both litmus strips? ____________________

     ____

     Note: These two changes, in this order, show the presence of chlorates.
   - (b) Squeeze about one inch of toothpaste into a a test tube. Half fill the test tube with water. Shake well. Allow the mixture to stand until the solids settle. Pour about one-fourth of the top liquid into another test tube and add 7 ml of concentrated hydrochloric acid. Record any change you observe. Dip a strip each of red and blue litmus paper in the solution. Record any change you observe.

DATA CHART

Toothpaste	Tumeric test for boron	Tests for chlorates	
		with HCl	with Litmus
Sample No. 1			
Sample No. 2			
Sample No. 3			

Questions to answer

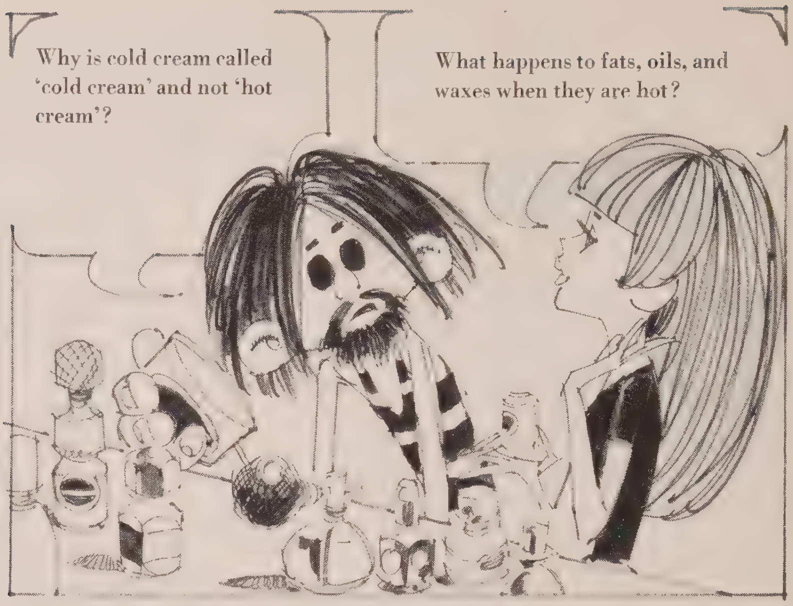
1. Where is boron found in the Periodic Table?

   ____
2. Borax is sodium tetraborate. Look up the formula and write it here. ____________________

   ____
3. What is the formula for potassium chlorate?

   ____
4. What chemical element is present in both borax and in potassium chlorate? ____________________

   ____


Why is cold cream called 'cold cream' and not 'hot cream'?
What happens to fats, oils, and waxes when they are hot?

Making Cold Cream

Facial creams, when properly used, help your personal appearance. Three principal benefits may come from their use: (1) softening of the skin, (2) cleaning of the skin, and (3) protection of the skin against the drying action of wind and sun. Remember that the skin is a living tissue and its nutrition comes from within the body.

What is cold cream? Cosmetic creams, sold as "cold cream", "skin cream", "beauty cream", and others, have many likenesses. They are ordinarily emulsions of fats, oils, and waxes in water. Beeswax and spermaceti are the waxes most commonly used, although paraffin wax is sometimes included. Oils most commonly used are colorless mineral oil, oil of sweet almond, and olive oil. Lanolin, a fat obtained from cleaning wool, is contained in many cold creams.

A small amount of borax solution is added to keep the waxes mixed with water. The borax and beeswax both act as emulsifying agents. The old-fashioned formula followed by druggists many years ago called for rose water. Today, plain water is used for the borax solution. A selected perfume is stirred in as the cream cools to near room temperature.

Experiment 19.3

Equipment and Supplies per Lab Station

beaker—150 ml; beaker—250 ml; evaporating dish
stirring rod, glass; test tube; thermometer
burner; clamp for test tube; clamp for thermometer
ring stand with ring and wire gauze; beaker tongs
2.5 g synthetic spermaceti wax; 2.4 g white beeswax
12 g white mineral oil, colorless; 4 ml borax solution

What to do

1. In your evaporating dish place the following quantities of waxes which have been cut into small pieces:
   2.5 g synthetic spermaceti wax
   2.4 g white beeswax
2. Place the evaporating dish on the top of a beaker of boiling water, as illustrated. Heat until the waxes are just melted at 80°C.
3. Pour 12 g of mineral oil *slowly* into the completely melted waxes. *Stir constantly.*
4. Turn out the fire. Remove the beaker and evaporating dish from the ring stand. Use beaker tongs or a very thick pad of paper towel to protect your hands.
5. Heat 4 ml of borax solution in a test tube to just 80°C. See Fig. 97.
6. Add the hot borax solution *slowly* to the melted wax and oil. *Stir the mixture constantly.* Constant stirring is necessary to obtain a good emulsion.
7. Remove the evaporating dish from the top of the hot beaker. Continue to stir the cold cream until it has cooled to room temperature. A drop of perfume may be added when the cold cream is at about 40°C. Stir in thoroughly.
   Transfer the product to a suitable container.
8. Describe the cold cream.
   (a) What color is it? ____________
   (b) What is its consistency, that is, thick?, soft?, smooth? ____________
   (c) Does a tiny bit spread smoothly on the skin of your wrist? ____________
   Note: Be sure to wash it off completely.

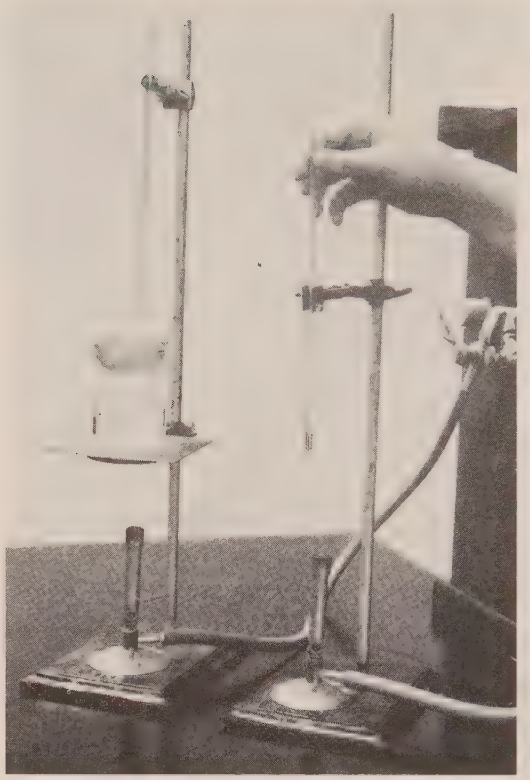


Fig. 97. *(Clara Aich and George Senty)*

Questions to answer

1. How do the emulsifying agents act to keep the waxes and oil mixed with the water? (Remember the egg white with the oil and water in Experiment 16.4.) ____________

   ____

   ____
2. Why is the perfume not added until the cream is cooled nearly to room temperature? ____________

   ____
3. What is the source of spermaceti wax? (*Hint*: Look up *spermaceti* in the dictionary.) ____________

   ____

   ____

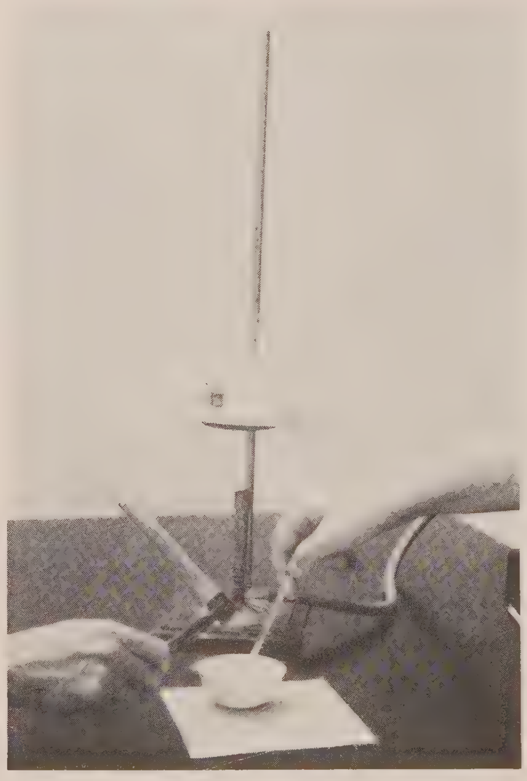


Fig. 98. *(Clara Aich and George Senty)*

Making Lipstick

Lipsticks are similar to cold cream in composition, but are firmer in consistency. Colors vary widely. They may be flavored.

Experiment 19.4

Equipment and Supplies per Lab Station

beaker—150 ml; beaker—250 ml; evaporating dish
glass stirring rod; test tube, 13 mm × 100 mm; burner
ring stand with ring and wire gauze
masking tape; aluminum foil, 8 cm × 6 cm
3.0 g synthetic spermaceti wax; 2.6 g white beeswax
4.5 g white mineral oil; 1 or 2 drops of methylene blue

What to do

1. Melt together in your evaporating dish over hot water, as you did for cold cream:
   3.0 g synthetic spermaceti wax
   2.6 g white beeswax
2. Add *slowly*, with constant stirring, 4.5 g of white mineral oil to the melted waxes. Stir until well blended. Add one or two drops of any dye solution your teacher may provide.
3. Roll a strip of aluminum foil 8 cm × 6 cm around a very small test tube. Fasten the aluminum with masking tape. Crimp one end over to close it. See Fig. 99-B.
4. Push the warm lipstick mixture into the aluminum holder. Set the holder in a beaker of warm water, 60°C, so that the lipstick mixture will melt down into the holder. See Fig. 99-D.
5. Harden the lipstick by placing the holder in a beaker of ice water. Peel back the aluminum foil. Form a point on the lipstick. Try it on your wrist. Wash it off carefully.

Questions to answer

1. Why could the borax be omitted from the lipstick? ______________________

   ____
2. Look up in reference books kinds of dyes used to color lipstick.
   (a) What are some of their sources? ____________

   ____

   (b) Why has the use of some of these dyes been prohibited by law? ______________

   ____

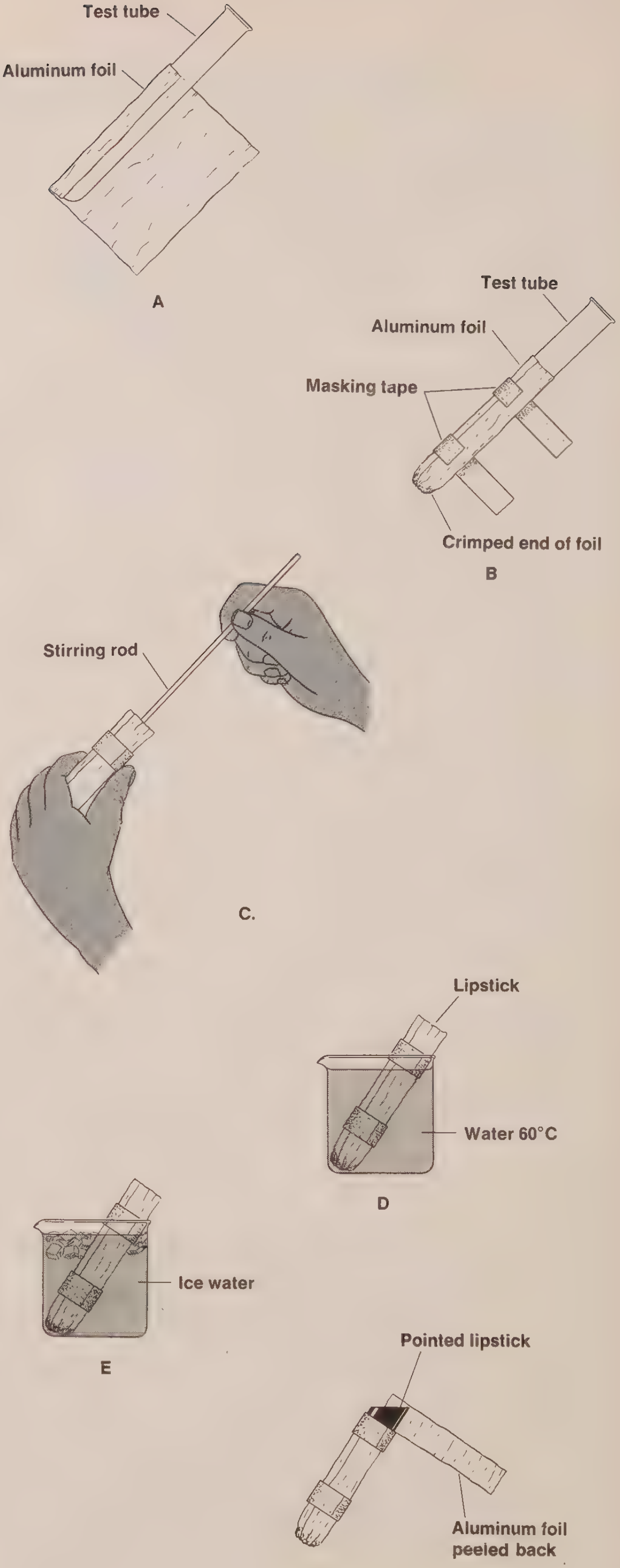


Fig. 99.

Laboratory Experiments in

ACTION CHEMISTRY

UNIT VIII

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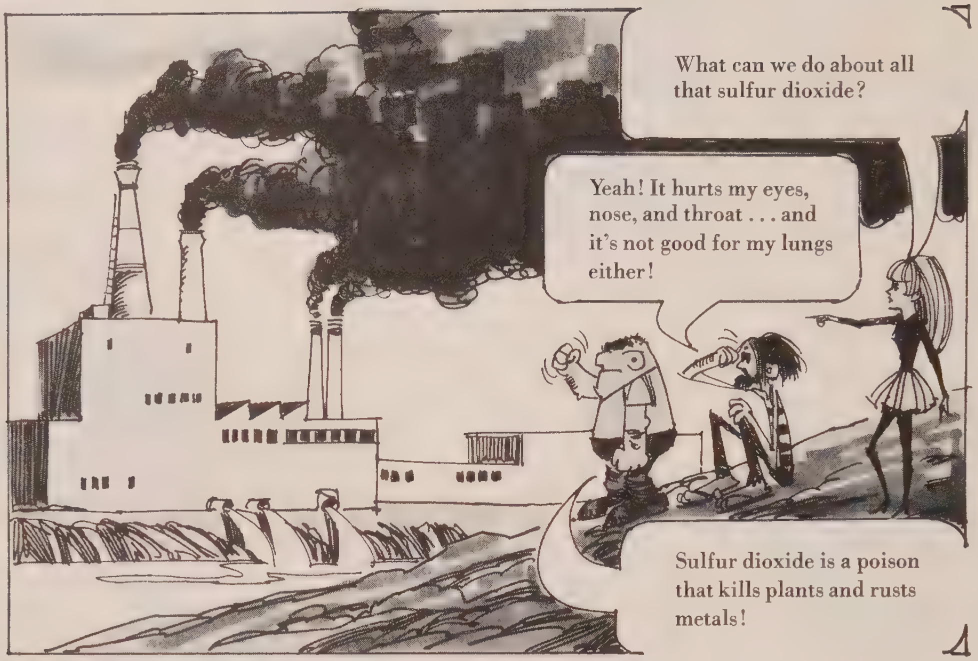
Chapter 20


What can we do about all that sulfur dioxide?
Yeah! It hurts my eyes, nose, and throat . . . and it's not good for my lungs either!
Sulfur dioxide is a poison that kills plants and rusts metals!

The REEKING Air

In this experiment you will see some of the effects of air pollutants.

Since the pollutants are harmful to your health, do the laboratory work under the hood. To keep the pollution to a minimum, this should be done as a group experiment or as a demonstration.

Experiment 20.1

Equipment and Supply per Lab Station

PART I
bottle—large mouth with stopper or cap
glass plate; stirring rod; (2) watch glasses
burner; combustion spoon
calcium carbonate—marble chips
blue litmus paper
magnesium ribbon—about 2-cm long
sulfur; water

PART II
(2) bottles—must be dry; (2) glass plates
ammonia water concentrated
hydrochloric acid—concentrated
perfume or oil of wintergreen

PART III
bottle; glass plate
graduated cylinder; tongs
camphor; water

PART I

Sulfur dioxide is a pollutant formed during the burning of most fuels.

What to do

1. Pour about 25 ml of very cold water into a 500-ml wide-mouthed bottle. Cover the top of the bottle with a glass plate.
2. Place a piece of sulfur about the size of a small pea in a combustion spoon.
3. Light a burner. Heat the sulfur until it starts to burn.
4. Lift the glass plate. Insert the spoon into the bottle. Hold the spoon above the water. Put the glass plate back over the mouth of the bottle as far as you can. Keep it there while the sulfur is burning. See Fig. 100. CAUTION: DO NOT INHALE THE GAS PRODUCED. What gas is formed when sulfur burns in air (oxygen)? ______

______

5. Remove the spoon when the sulfur stops burning. Stopper the bottle quickly. Shake it vigorously for two to three minutes.
6. Dip a clean stirring rod into the sulfur dioxide water. Touch the wet stirring rod to a piece of blue litmus paper.
   (a) What is the color change of the litmus?

   ______

   (b) What type of compound turns litmus this color? ______

   (c) What compound could be formed when the gas from burning sulfur dissolves in water?

   ______

   Complete the equation for the reaction.

   sulfur dioxide + water → ______

   $SO_2 + H_2O \rightarrow$ ______

7. Place a few marble chips on a watch glass. Pour about 10 to 15 ml of the sulfur dioxide water onto the chips. Observe the chips for five minutes. Write down anything you observe.

   ______

   If a gas escapes, what do you think it might be?

   ______

8. Place a piece of magnesium ribbon on a second watch glass. Cover the magnesium with the sulfur dioxide water. Observe the metal five to ten minutes. Write down your observations. ______

   ______

   If a gas escapes, what do you think it is? ______

   ______

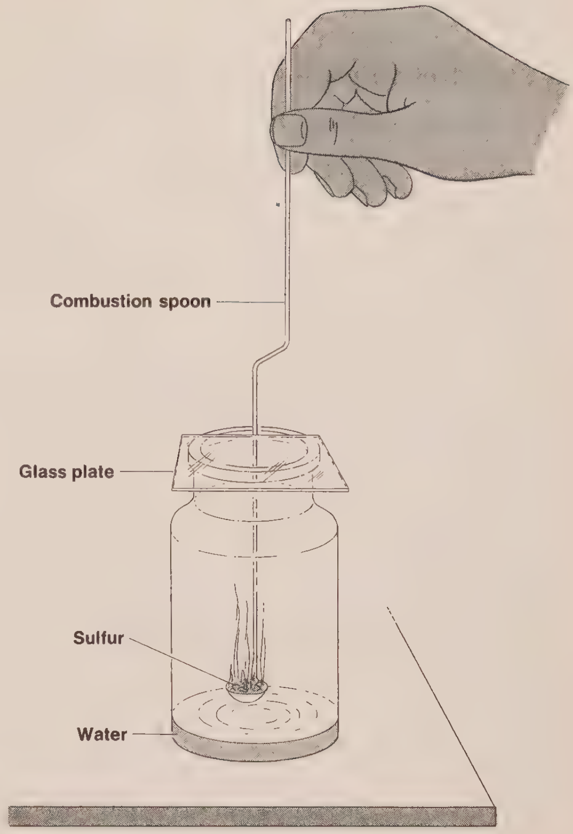


Fig. 100.

Questions to answer

1. Complete the following equation for the burning of sulfur in air.

   sulfur + oxygen → ______

   $S_8 + 8\,O_2 \rightarrow 8$ ______

2. (a) Complete the following equation for the reaction in step 7.

   sulfurous acid + marble chips → ______ + ______ + ______

   $H_2SO_3 + CaCO_3 \rightarrow$ ______ + ______ + ______

   (b) What effect would air polluted with sulfur dioxide have on buildings made of limestone or marble? ______

   ______

3. (a) Complete the following equation for the reaction in step 8.

   sulfurous acid + magnesium → ______ + ______ + ______

   $H_2SO_3 + Mg \rightarrow$ ______ + ______ + ______

   (b) What would be the long range effect of air containing moisture and sulfur dioxide on metal structures such as bridges? ______

   ______

PART II

A volatile substance is one which changes to a gas at a low temperature. Some examples of volatile substances are perfumes, concentrated hydrochloric acid, and household ammonia.

What to do

1. Open a bottle of perfume or oil of wintergreen. Let it stand unstoppered for a few minutes. Ask your classmates to say when they can smell the volatile substance wherever they are in the room. If any odor is detected at a distance from the bottle, explain how this is possible. ______

   ______

2. Place two or three drops of concentrated hydrochloric acid in a dry bottle. Cover the bottle immediately with a glass plate. Write down any observations. ______

   ______

3. Pour about 1 ml of strong household ammonia into another dry bottle. Cover the bottle immediately with a glass plate. Write down any observations. ______

   ______

4. Place the bottles next to each other. Remove the covers carefully. Record all you observe. ______

   ______

Questions to answer

1. Complete the following equation for the reaction in step 4.

hydrochloric acid + ammonia water → ______________ + ______________

HCl + NH_4OH → ______________ + ______________

2. What are two things that can happen to a gas released into the air?

(a) ____

(b) ____

PART III

Incomplete burning of substances often produces undesirable solids. Such solids may float about in the air as pollutants and finally get into our lungs.

What to do

1. Obtain a piece of camphor no larger than a small pea.
2. Hold the camphor with tongs. Ignite it carefully by placing it in a flame.
3. Drop the burning camphor into a wide-mouthed bottle. Cover the bottle quickly with a glass plate. What do you see happening? ____________

    ____

4. Allow the camphor to burn out. Add about 25 ml of water to the bottle. Shake the bottle so that the water hits the sides.

    (a) What happens? ______________________

    (b) Describe the solid product. ____________

    (c) What do you think the solid product might be? ______________________

Questions to answer

1. Fires in industrial plants may produce soot. What happens to this soot? ______________

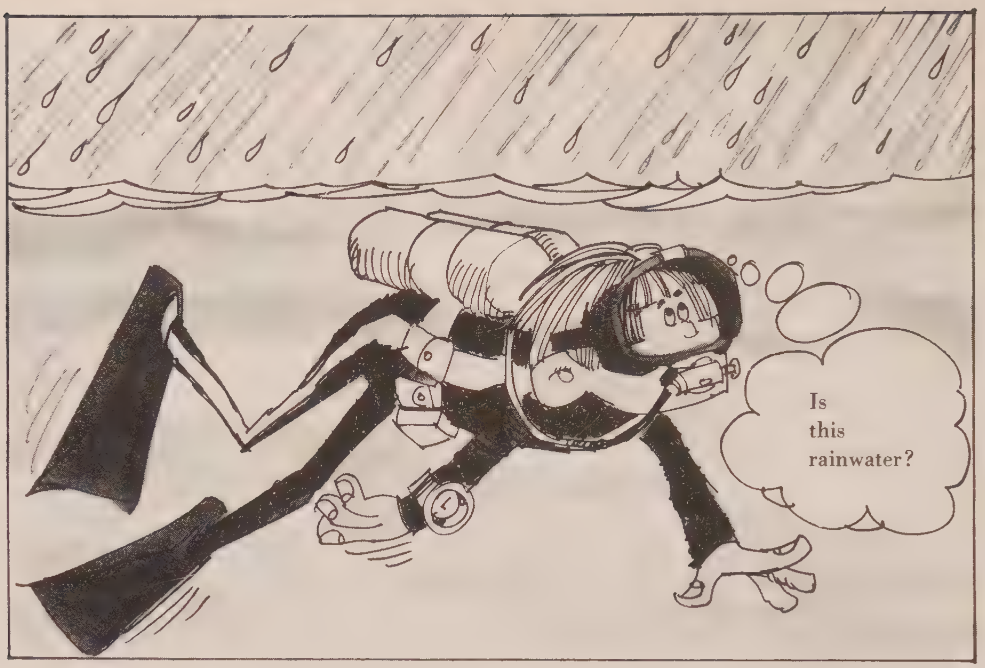
    ____

2. Why do you think many communities have ordinances banning home incinerators and the burning of trash and/or leaves? ______________

    ____

3. Why do you think the control of air pollution might be a national rather than a local problem? ______________________

    ____


Is this rainwater?

Another Detective Assignment—
IDENTIFY THE RAINWATER

Can you tell which of the samples given you is rainwater? Perform some tests to get some clues which will help you form an answer.

Experiment 20.2

Equipment and Supplies per Lab Station
graduated cylinder—10, 25, or 50 ml
stirring rod; (3) test tubes; test tube rack
labels or a marking pen
(3) vials pH indicator test papers
(4) water samples to be tested
5 ml barium chloride solution
5 ml silver nitrate solution

What to do

1. Obtain about 50 ml of a water sample to be tested. Observe its appearance and odor. Write your observations in the Data Chart.

2. Number three test tubes consecutively. Pour 10 ml of the sample to be tested into each test tube.

3. Pour 5 ml of barium chloride solution into Test Tube No. 1. Note any reaction which seems to occur. Write any observations in the Data Chart.
4. Pour 5 ml of silver nitrate solution into Test Tube No. 2. Write any observations in the Data Chart.
5. Dip a clean, dry, stirring rod into the third test tube. Touch the wet rod to a strip of pH indicator paper. After comparing the color of the paper with those on the pH scale, write the pH in the Data Chart.
6. Evaporate 10 ml of the sample on a watch glass placed on a beaker of boiling water. See Fig. 101 for the apparatus setup. Add any observations to the Data Chart.
7. Repeat steps 1–6 with each of the other samples given you. Record all observations in the Data Chart.

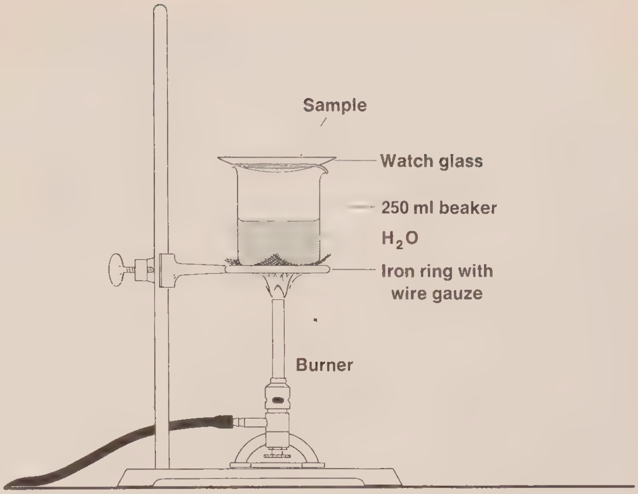


Fig. 101.

DATA CHART

Sample No.	Appearance	Odor	Add solution $BaCl_2$	Add solution $AgNO_3$	pH	After evaporation any residue?

Questions to answer

1. If a precipitate forms with barium chloride, what might it be? (The solubility chart can give you some clues.) ______________________
____
2. What precipitate might be formed upon the addition of silver nitrate? ______________________
____
3. Do you think that rainwater will form a precipitate with either barium chloride or silver nitrate? ______________ Why or why not? ______
____
4. As rain falls, carbon dioxide tends to dissolve in the rainwater. Does this make the water slightly acid or basic? ______________________
5. What appearance does rainwater have? ________
____
6. Which sample(s) do you think might be rainwater? ______________ Why? __________
____

CLEARING THE WATER

Hey, man! Coagulate your turbidity into floc!

Untreated water supplies contain suspended material which does not settle quickly. Such water is said to be *TURBID, which means cloudy*. When turbid water is treated with certain chemicals, suspended material settles more quickly.

The chemicals used to increase the settling rate are called COAGULANTS. They form a sticky substance called a FLOC. Suspended particles stick to the floc. Eventually enough particles stick to the floc to form large clumps. The clumps settle to the bottom.

The amount of coagulant depends upon the amount and kind of suspended matter. How much coagulant to use must be determined for each sample of turbid water.

In this experiment, find which amount of coagulant produces good floc formation and causes the fastest settling.

The coagulant you will use is a solution containing 1 g of aluminum sulfate in one liter of solution. Good floc formation may not mean the fastest settling rate. If you watch closely and follow directions carefully you will get an accurate answer.

Experiment 20.3

Equipment and Supplies per Lab Station

(7) beakers—1-liter or 400 ml; clock
(7) graduated cylinders—100 ml
(6) pipets and bulbs or burets—25 ml
stirrer; turbid water sample
aluminum sulfate solution, 1,000 mg/liter

What to do

This is a group experiment. Each team will carry out steps 2 through 5 for the beaker and sample assigned him by the teacher. It is important that all team members do the same steps at the same time! The student in charge of Beaker No. 1 (the control) can act as timekeeper.

1. Set up seven one liter beakers numbered 1 through 7 respectively, for each group. The beaker number corresponds to the assigned number. Your assigned number is ____________
   Note: You may be instructed to use 400-ml beakers instead of liter beakers.
2. Pour 400 ml of the turbid water into your one liter beaker. *Note*: Pour only 200 ml of the turbid water into a 400-ml beaker.

3. Use a pipet to add the assigned volume of the aluminum sulfate solution to your beaker—all members of the team at the same time. *Note:* You will be using either a 400-ml sample or a 200-ml sample, but not both. The amounts of aluminum sulfate for both quantities are given in the following chart.
4. Stir the liquid in the beaker vigorously for two minutes. Then stir the liquid *very slowly* (10 to 15 revolutions per minute) for eight minutes.
5. Set the beakers aside to settle for 20 minutes. *Do not* disturb the samples of water as you observe them during the 20 minutes.
6. Inspect all seven beakers with the other members of your team during and at the end of the 20 minutes settling time. Describe the samples as follows:
   (a) Which sample formed the floc the fastest? ______
   (b) Which sample seems to have the most floc? ______
   (c) Which sample seems to form the largest particles? ______
   (d) Which sample seems to have the least floc? ______
   (e) Which sample seems to settle best? ______
   (This one will have the most clear water above the settling floc.)
7. Discuss your observations with your team members. What amount of coagulant per 400 ml (or per 200 ml) would you recommend for obtaining the clearest sample of water in 20 minutes? ______

Beaker No. Assignment No.	ml of aluminum sulfate solution added to the turbid water		Observations
	to 400 ml	to 200 ml	
1	none (control)	none (control)	
2	3.3	1.6	
3	6.6	3.3	
4	10.0	5.0	
5	13.3	6.6	
6	16.6	8.3	
7	20.0	10.0	

Questions to answer

1. Where are chemical coagulation processes most likely to be used in your community? ______
2. Why isn't filtering used to obtain clear water from turbid samples? ______

Extra for experts

What other type of coagulant may be used to clear up turbid water? ______

(The answer may be found through library research or from an interview with the operator of the water treatment plant in your community.)

pH AND THE FLOC FORMATION

The pH of a turbid water sample depends upon the source of the water. Water from a river may have a different pH than a sample collected from a puddle on the street.

Does the pH of the water sample affect the floc formation?

Experiment 20.3-A Extension

Equipment and Supplies per Lab Station

(2) beakers—1 liter or 400 ml; graduated cylinder—100 ml
pipet and bulb for dispensing coagulant
(2) stirring rods; clock
pH indicator test papers
aluminum sulfate solution, 1,000 mg/liter
hydrochloric acid, 1*M* in dropping bottle
sodium hydroxide, 1*M* in dropping bottle
turbid water sample from Experiment 20.3

What to do

1. Obtain 800 ml of the turbid water sample used in Experiment 20.3. Pour 400 ml of the water into each of 2 one-liter beakers. Number the beakers 1 and 2 respectively.
   Note: You may be instructed to use 400-ml beakers and place 200 ml of the turbid water in each.

2. Add 10 drops of 1 *M* hydrochloric acid to Beaker No. 1. Add 10 drops of 1 *M* sodium hydroxide to Beaker No. 2.

3. Dip a clean, dry, stirring rod into the water in Beaker No. 1. Touch the wet stirring rod to a strip of dry pH test paper. Immediately compare the color of the test paper with the color chart on the test paper vial.
   What is the pH of this turbid water sample?

   ____

Fig. 102. Two students are stirring two samples of turbid water. *(Clara Aich and George Senty)*

4. Repeat step 3 using the water in Beaker No. 2.

   What is the pH of this sample of water? ______

   ____

5. To each beaker, add the amount of aluminum sulfate your team recommended as giving the best floc formation and settling rate in Experiment 20.3.

6. Stir both beakers vigorously at the same time for two minutes. Then continue to stir very slowly at 10 to 15 revolutions per minute for eight more minutes.

7. As the beakers stand undisturbed for the 20 minutes as in Experiment 20.3, observe the floc formation and settling in both beakers. Write all observations below.

   Beaker No. 1 pH ______ ____________

   ____

   Beaker No. 2 pH ______ ____________

   ____

Questions to answer

Do you think that pH affects the floc formation?

____

Why do you think so? ____________________

____

Extra for experts

What other factors do you think might affect the floc formation and settling rate? (Go to the library for additional information.) ______________

____

____

____

Chapter 21

Hey, how come my plant is so small compared to yours?
Maybe you have too much sand in your loam!
Naaa, just too many rocks between the ears!

Some Properties of **SOIL**

Plants grow well in loam soil. Loam is a mixture of four main soil types: gravel, sand, clay, and silt. Gravel and sand are coarse and porous. Clay is very fine grained. Silt is also fine grained, but not as fine grained as clay.

Soils vary widely in different areas. The relative amounts of sand, gravel, clay, and silt differ. The amount of moisture the soil will hold depends on the mixture of soil types. The pH and nitrogen content may differ.

What are some kinds of soil in your area?

Experiment 21.1

Equipment and Supplies per Lab Station — for team of four students

(4) gas-collecting bottles
(5) graduated cylinders; (2-4) test tubes
(2-4) burners; flashlight (1 or 2 will supply whole class)
(2-4) ring stands with clamps; stopper to fit graduated cylinders
copper wire-22 gauge, about 3 inches; litmus paper, red and blue
sample of loam soil; (3) samples of different local soils
distilled water

What to do

For both parts of this experiment students should work in teams of four. One sample of loam soil and three samples of soils from different locations in your area will be provided. Each student will make the tests on one of the samples. The team will compare observations. All members record results on all four samples.

PART I

First day

1. Label a 50-ml graduated cylinder according to the sample assigned you.
2. Fill the cylinder to the 35-ml mark with the soil sample. Add distilled water to the 50-ml mark.

3. Stopper the cylinder tightly. Shake vigorously for one to two minutes.
4. Place the cylinders from each group where they will not be disturbed. Allow them to stand overnight.

Second day

5. Record the milliliters of each distinct layer of soil in your sample and in each of the samples in your group, in Data Chart 1.
6. Test the liquid above the soil with both red and blue litmus paper. Record whether it is acid or base, in Data Chart 1.
7. Examine the liquid above the soil.
   (a) Is it clear or cloudy? ____________________
   (b) Fill a second 50-ml graduated cylinder with water. Shine the beam of a flashlight so that the light passes first through the water and then through the liquid above the soil. Record the result in Data Chart 1.

   What do you see? ____________________
   How do you explain this? (Review Experiment 16.3.)

   ____

   ____

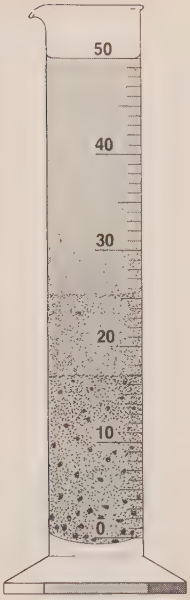


Fig. 103. Distinct layers of soil in the cylinder on the second day.

DATA CHART 1

Soil Sample	Coarse layer(ml)	Medium layer(ml)	Fine layer(ml)	Litmus test	Colloid Yes or No
1. Loam					
2.					
3.					
4.					

PART II

The same team of four students should do these tests as they did the tests in PART I. Use samples of the same soils tested in PART I. The soil samples should be as dry as most surface soil is in sunny dry weather. They should not be damp as after a rainstorm or after irrigation.

What to do

1. Test for nitrogen as in Experiment 16.5, PART I.
   (a) Place about one-half teaspoonful of your soil sample in a clean, dry test tube.
   (b) Suspend a strip of damp red litmus paper by means of a piece of copper wire so that the litmus paper hangs down in the top of the test tube.

(c) Attach the test tube to a ring stand. Light a burner. Heat the test tube strongly for two to five minutes.
What change, if any, do you see in the

litmus paper? ________________________

__________ Explain. ________________________

____

Record the result for the test for nitrogen in Data Chart 2.

2. Label a wide-mouthed bottle, such as a gas-collecting bottle, for your sample of soil. Fill the bottle about two-thirds full with the soil.

3. Fill a 50-ml graduated cylinder to the 50-ml mark with water. Add the water a little at a time. See how much water you can add without having any rise above the soil. This will be the volume absorbed.

   To calculate the volume absorbed, subtract the volume remaining in the graduated cylinder from the 50 ml at the start. Record the milliliters absorbed in Data Chart 2.

4. Put the uncovered bottles from each group in a safe place. Look at them every day until the soil samples are dry. Record in Data Chart 2 the number of days required to dry each sample:

DATA CHART 2

Soil Sample	Nitrogen (Yes or No)	Water absorbed (ml)	Number of days until dry
1. Loam			
2.			
3.			
4.			

Questions to answer

1. What differences did you find in the amounts of coarse, medium, and fine material in the layers in the different samples tested? ________________________

____

____

2. How do your local samples compare with the loam sample? ________________________

____

3. Did the samples vary in reaction to litmus?

____

4. Did the liquid from above all the soil samples show a colloidal property? Explain. ____________

____

5. Did you find nitrogen in all samples? ____________

____

6. Do you see a relationship between the absorption of water, the drying time, and types of soil particles found in the layers? Explain.

____

____

____

Soil pH

The pH of soil is important. Only a few plants will grow well in a highly acid soil. Almost none will grow in a basic soil. Most food crops will grow in soils over a range of pH 6.0 to pH 7.5. For example:

wheat	pH 6.0 to 7.0	green peas	pH 6.0 to 7.5
potatoes	pH 5.5 to 7.0	oranges	pH 6.0

Cranberries are different. They grow well in bogs at pH 4.0.

Sometimes it is necessary to neutralize an acid soil in order to grow a special crop. Lime, which is calcium hydroxide, may be added to the soil for this purpose.

Highly basic soils are reclaimed by LEACHING out the basic salts. The land to be reclaimed is flooded frequently with water over a period of a year. The basic salts are thus *LEACHED (washed out)* of the surface soil. They settle down into the subsoil.

In this experiment you will neutralize a sample of acid soil.

How do you "reclaim" this land?
Yeah, how do you make this land bloom again?
Yeah, it has a pH of 8?

Experiment 21.1-A Extension

Equipment and Supplies per Lab Station

beaker-150 ml; dropper with long tip; graduated cylinder-25 ml; balance; stirring rod
distilled water; limewater
litmus paper, red and blue
Universal Indicator paper
sample of acid soil

What to do

1. Weigh a 150-ml beaker. Weigh 1 g of a soil sample in this beaker.
2. Add 10 ml of distilled water to the sample. Stir to get an even suspension of the soil in the water.
3. Dip a strip each of red litmus, blue litmus, and Universal Indicator paper in the soil mixture. Record the reactions.

   red litmus ______________________

   blue litmus ______________________

   Universal Indicator – pH ______________________
4. Place 15 ml of limewater in a graduated cylinder. Stir the limewater, drop by drop, into the soil mixture until it is just neutral to the litmus and to the Universal Indicator. Calculate the number of milliliters required to neutralize the one gram sample of soil.

15 ml – ________ml remaining in the cylinder =

________ml required to neutralize 1 g of soil

Question to answer

Inquire from an agriculture experiment station how the pH of soil is tested in the field. __________

____

____

How do you know when you've put enough water in them?

Moisture Content of SOIL

Determine the moisture content of one of the samples of soil which have been provided.

Obtain the moisture content of each of the other samples as determined by other members of your class.

Experiment 21.2

Equipment and Supplies per Lab Station

beaker-250 ml; evaporating dish
funnel; graduated cylinder
balance, laboratory; drying oven
ring stand with small ring
filter paper; (4) different soil samples per class
distilled water

What to do

PART I

1. Weigh your evaporating dish. Record the mass in the Data Chart.
2. Add about 5 g of one of the soil samples. Weigh carefully. Record the exact mass of the evaporating dish plus the soil. Calculate the mass of the soil sample. Record in the Data Chart.
3. Place the evaporating dish with the soil sample in a low-temperature drying oven over night. Again weigh carefully and record. *Save the dry soil for PART II of this experiment.*
4. In the Data Chart calculate the mass of the soil after drying.
5. In the Data Chart calculate the mass of the water lost by drying. (Subtract the mass of the dry soil from the mass of the soil before drying. The difference equals the mass of water lost.)
6. Calculate the percentage of the original mass of the soil that was water. Record in the Data Chart.

$$\frac{\text{mass lost}}{\text{original mass}} \times 100 = \text{percentage of original mass that was water}$$

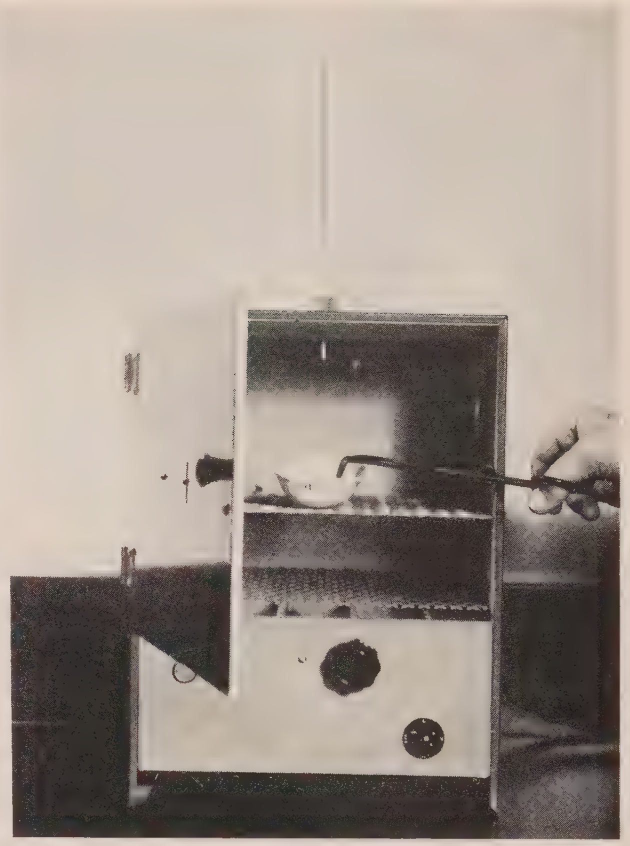


Fig. 104. Placing the soil sample in the drying oven. *(Clara Aich and George Senty)*

DATA CHART

	Masses Weighed (grams)	Masses Calculated (grams)
Evaporating dish		X
Evaporating dish plus soil (*Before* drying)		X
Soil *before* drying – original mass (Subtract mass of dish)	X	
Evaporating dish plus soil (*After* drying)		X
Soil *after* drying (Subtract mass of dish)	X	
Mass lost in drying (Mass of soil before drying minus mass of soil after drying)	X	
Percentage of original mass that was water	X	

7. Obtain the percentage of water in other samples from your classmates.

Sample	Percentage of water	Sample	Percentage of water
1		3	
2		4	

PART II

1. Transfer the dry soil from PART I to a filter apparatus. Be careful not to lose any of the soil.

2. Measure 15 ml of water in a graduated cylinder.

3. Pour the water from the cylinder onto the soil a little at a time until the soil is wet and water begins to run through the funnel.

4. Catch the water that comes through the funnel in a beaker. Return it to the graduated cylinder.

5. Calculate how many milliliters of water were held by the soil.

   15 ml − ml returned to and left in cylinder

   = ml held by the soil

6. Calculate the milliliters of water held by each gram of dry soil. Use the mass of the dry soil found in PART I.

   $$\frac{\text{milliliters of water held}}{\text{mass in grams of dry soil}} = \begin{array}{l}\text{ml of water per}\\ \text{gram of soil}\end{array}$$

7. Obtain the milliliters of water per gram of other soil samples.

Sample	ml of water/g of soil	Sample	ml of water/g of soil
1		3	
2		4	

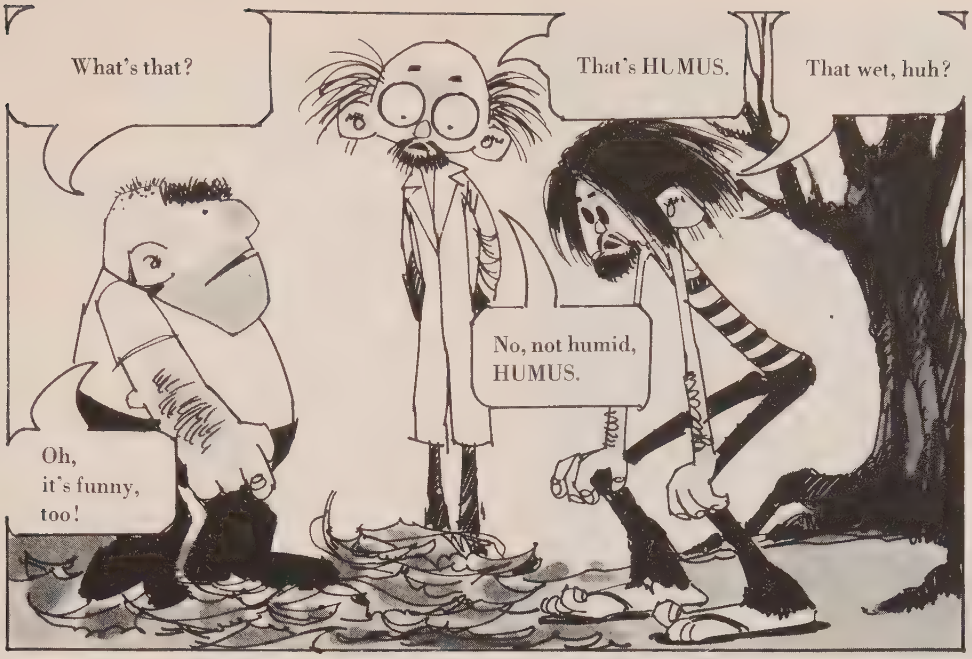
Questions to answer

1. What differences do you see in the moisture content of the soil samples? ____________________

   ____

2. How do you explain these differences? __________

   ____


What's that?
That's HUMUS.
That wet, huh?
No, not humid, HUMUS.
Oh, it's funny, too!

Humus in SOIL

Plants grow, shed leaves and blossoms, die and fall to the ground. Dead plant material decays and forms *humus*. Humus mixes with the soil and enriches it by replacing nutrients taken out of the soil by growing plants.

In this experiment you will determine the percentage of humus in a sample of soil or in peat moss.

Experiment 21.3

Equipment and Supplies per Lab Station

crucible; balance, laboratory burner; clay triangle; drying oven ring stand with small ring sample of soil or peat moss

What to do

First day

1. Weigh a crucible. The mass is ______________g
2. Fill the crucible about one fourth to one third full of soil or peat moss.
3. Dry the sample overnight in a drying oven set at about 50°C.

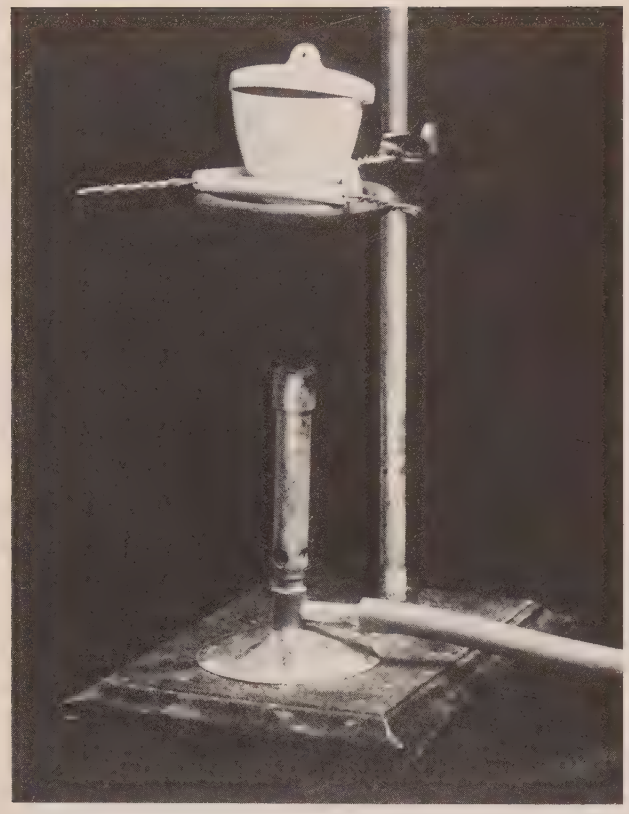


Fig. 106. *(Clara Aich and George Senty)*

Second day

4. Examine the dried sample. Describe its appearance. ______________________________
____
5. Weigh the crucible and dry sample. Record the mass and calculate the mass of the dry sample.
   (a) Crucible + dry sample ____________g
   (b) Empty crucible (step 1) ____________g
   (c) Mass of dry sample (a − b) ____________g
6. Place the crucible in a clay triangle on a ring stand. Heat with a hot flame. See Fig. 106.
7. Stir the sample frequently. Continue the heating for about 30 minutes. Keep the flame as hot as possible.
8. Cool the crucible and contents. Has there been any change in the appearance of the sample? Describe the sample. ______________________________
____
____
9. Weigh the cool crucible and contents. Record the mass. Calculate any loss of mass.
   (a) Crucible + dry sample (step 5) ________g
   (b) Crucible + burned sample ________g
   (c) Loss of mass (a − b) ________g
10. Assume that any loss of mass was due to the burning of humus present in the sample. Calculate the percent of humus.

$$\frac{\text{mass lost } \_\_\_\_\_\_\_\_\_\_ \text{g}}{\text{mass of dry sample } \_\_\_\_\_\_\_\_\_\_ \text{g}} \times 100 =$$

______________% humus in dry sample

Note: You calculated the mass of the dry sample in step 5.

Questions to answer

1. Why do you think it is possible to burn humus out of a soil sample? (*Hint:* Do starches, fats and proteins burn?) ______________________________
____
____
2. The *ash* is the material left after all burnable substances have been burned away. What materials would you expect to find in the ash from this experiment? ______________________________
____
____

The Color In Leaves

Green is the color you notice most about plants. *The green coloring material is called CHLOROPHYLL.* Without chlorophyll plants could not make food.

Yellow color is also found in the leaves of plants. This is due to CAROTOIDS.

The colored substances in the leaves, known as PIGMENTS, can be separated by paper chromatography. This is the process we shall use in this experiment.

Experiment 21.4

So, what is chlorophyll?
Oh, that's the stuff in toothpaste.
No. Chlorophyll makes plants green.
So, what's chlorophyll?

Equipment and Supplies per Lab Station

beaker—1 liter or 600 ml; evaporating dish; graduated cylinder-25 ml
mortar and pestle; scissors
10 ml acetone
filter paper-Whatman #1, 11 cm diameter
samples of leaves

What to do

1. Choose a few leaves from the samples supplied. From what plant did the leaves come? ______________ Are the leaves dark green, a lighter green, or a patterned type? ____

   ____

2. Tear the leaves into small pieces. Place the pieces in a mortar. Grind them to a paste with the pestle.

3. Add 8 to 10 ml of acetone. Continue grinding until the liquid is deep green color. CAUTION: MAKE SURE THAT NO FIRES ARE LIT WHILE ACETONE IS BEING USED.

4. Pour the green liquid into your evaporating dish.

5. Obtain a piece of Whatman #1 filter paper, 11 cm in diameter. Cut a strip one-fourth inch wide half way across the filter paper. Use scissors. Leave the strip attached to the filter paper. Look at the filter paper in Fig. 107.

6. Place the filter paper over the evaporating dish containing the green liquid. Bend the cut strip down into the liquid. The strip will act as a wick to draw the liquid up into the filter paper. See Fig. 107.
7. Turn a large beaker upside down over the evaporating dish.
8. Watch the colored rings form in the filter paper for about 20 minutes.

   (a) What color appears first? ______________

   How many rings of this color are there? ____

   Measure the width of each ring in cm. ______

   ____

   (b) What color appears next? ______________

   How many rings of this color are there? ____

   Measure the width of each ring in cm. ______

   ____

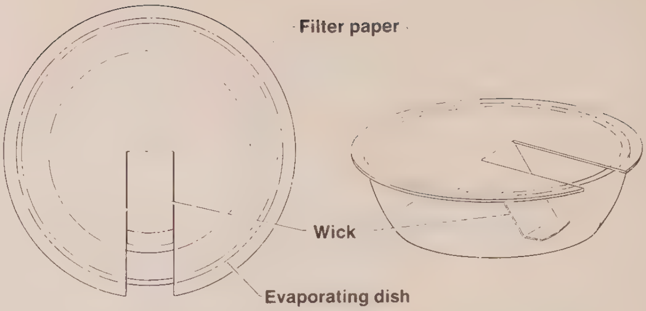


Fig. 107. The diagram at left is overview of strip in liquid. The diagram at right is sideview of the same strip.

9. Draw an outline of the filter and make a diagram of the rings you see after 20 minutes. Label each ring with the color and the name of the compound. Use the space provided.
10. Make a second diagram of the rings formed by the pigments from leaves different from those you tested. Record the name of the plant.

CHROMATOGRAPHS OF PIGMENTS EXTRACTED WITH ACETONE

Your Experiment	Another Experiment
Plant:	Plant:

Questions to answer

1. What differences did you see between the pigments in the leaves you tested and the pigments in the leaves of one other plant? ______

   ____

   ____

2. Look up the chemical reaction by which the plant manufactures sugars. Write the equation for the reaction in the space at the right.

3. What is the job of the chlorophyll in this reaction? ______________________

   ____

   Explain. ______________________________

   ____

UNFAIR COMPETITION

The farmer has a constant battle with plant diseases, insects, and weeds. Modern chemistry has given him weapons against these enemies. The trouble is that the new chemical weapons create as many problems as they solve. A very serious problem is pollution particularly of water resources.

Take a look at how this works. Fill in the following chart. Consult sources in the library. Your teacher will advise you. If possible, talk to people in charge at a local nursery or store supplying various pesticides.

Exercise

Questions to answer

1. Did you find some pesticides which do the job without endangering the workers who handle them, the consumers who use the products, or water resources? ____________________

   ____

2. Are there some pesticides whose advantages seem to outweigh fairly serious disadvantages?

   If so, what are they? ____________________

   List: *Advantages* *Disadvantages*

Advantages	*Disadvantages*
______________	______________
______________	______________
______________	______________
______________	______________

PESTICIDES

Type	Used to control	Commercial products	How applied	Dangers in handling	Pollution problems
Fungicides					
Insecticides					
Herbicides					

Green Plants at Work

The process by which green plants manufacture food is known as PHOTOSYNTHESIS. As you have learned, *SYNTHESIS means the chemical union of reacting substances to form new products.* The green plant, with the aid of chlorophyll, combines carbon dioxide from the air with water from the soil to make sugar. Sugar molecules combine to make starch. *PHOTO in the name of the process means that the energy needed comes from light.*

A little investigation will tell the story of how this works.

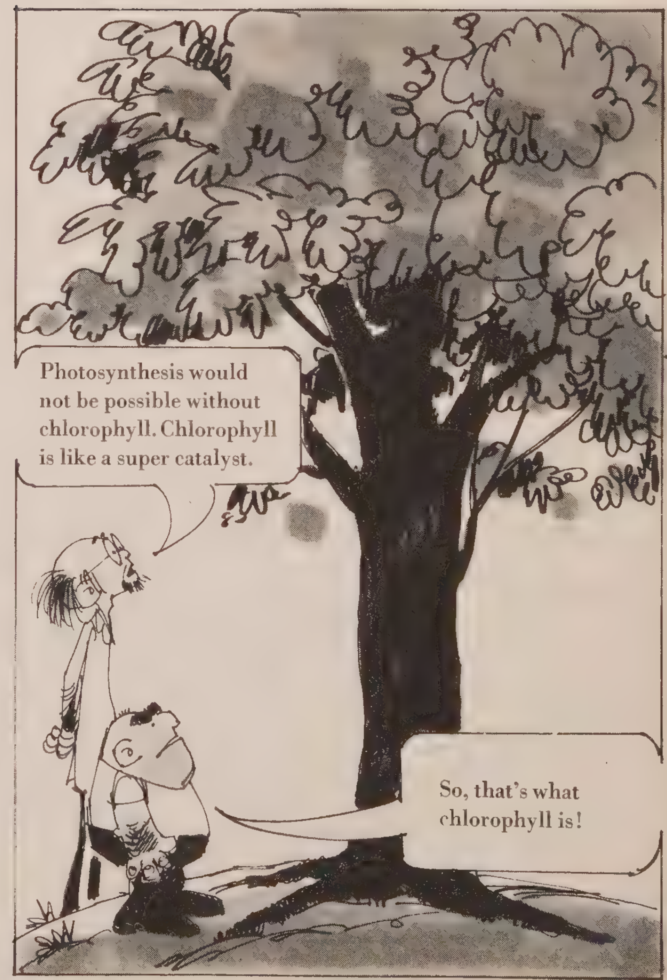
Experiment 21.5

Equipment and Supplies per Lab Station

beaker-400 ml; graduated cylinder; (5) test tubes, 18 mm × 150 mm; soda straws
goose-neck high intensity lamp
bromthymol blue indicator; distilled water
elodea water plant; sodium hydroxide, 1 *M*

What to do

1. Number five test tubes, 1, 2, 3, 4, and 5 respectively.
2. Place 20 ml of distilled water in each test tube.
3. Add 5 to 10 drops of bromthymol blue indicator solution. The water will turn a greenish color. This dye is green in neutral solutions, yellow in acid, and blue in base.
4. Test Tubes No. 1 and No. 2.
    - (a) Blow your breath through a soda straw into the water in each test tube in turn until the solution changes from green to yellow. This shows that an acid has been formed.
        - (1) What substance in the breath can react with water to form an acid? __________
        - (2) The chemical reaction for the formation of the acid is $CO_2 + H_2O \rightleftharpoons H_2CO_3$ (carbonic acid) The double arrows show that the carbonic acid can change back to CO_2 and H_2O.
    - (b) Keep Test Tube No. 1 for a control.
    - (c) Add a sprig of elodea plant to Test Tube No. 2.
    - (d) Set Test Tubes 1 and 2 aside in a safe place.


Photosynthesis would not be possible without chlorophyll. Chlorophyll is like a super catalyst.
So, that's what chlorophyll is!

5. Test Tubes No. 3, No. 4, and No. 5.
   (a) Keep Test Tube No. 3 as a control. It contains only water and bromthymol blue.
   (b) Test Tube No. 4 is also a control. Add one or two drops of 1 *M* sodium hydroxide. The color changes from green to the blue of the base.
   (c) Add a sprig of elodea to Test Tube No. 5.
6. Place all five test tubes from steps 4 and 5 in a large beaker. Put them under the strong light of a high intensity goose-neck lamp.
7. Watch the tests over a period of 30 minutes, Look for color changes. Note whether bubbles appear on the leaves of the elodea plant. Record your observations in the Data Chart.

DATA CHART

Test Tube	Color at start	pH at start	Observation after ½ hour	
			Color	Bubbles
1. Control (acid)				
2. With elodea				
3. Control (neutral)		pH 6-7.6		
4. Control (base)				
5. With elodea		pH 6-7.6		

Questions to answer

1. How would you explain any bubbles you may have observed on the leaves of the elodea plant?

____

____

2. How do you explain the color change in Test Tube No. 2? ______________________

____

____

3. Bromthymol blue is green over a pH range from pH 6.0 to pH 7.6. How does this fact help to explain any color change you may have observed in Test Tube No. 5? ______________

____

____

4. What evidences of photosynthesis have you observed in this experiment? ______________

____

Plant Nutrients

Soil is a mixture of chemicals. Plants ordinarily get the substances they need for growth from the soil and from the air. However, plants can be grown in "soil-less gardens" which are water solutions of essential chemicals. This method is known as HYDROPONICS.

One nutrient solution used in hydroponics contains the following chemical compounds:

ammonium phosphate, $NH_4H_2PO_4$
calcium nitrate, $Ca(NO_3)_2 \cdot 4\ H_2O$
magnesium sulfate, $MgSO_4 \cdot 7\ H_2O$
potassium nitrate, KNO_3

Chemical elements found only in traces in fertile soil can be added to the nutrient solution, or they may be omitted. This depends on the special requirements of the plant.

From what compounds does the plant obtain the chemical elements needed for growth? Are these compounds found in the air, in the soil or may they be found in the nutrient solution in a "soil-less garden?"

Exercise

What to do

Complete the following chart.

1. If the chemical element listed is present in one of the compounds in the hydroponic solution, write the formula for the compound in the hydroponic column and on the horizontal line with the element.
2. If the plant gets the element from a substance in the air, put a check mark in the column under "Air."
3. If the plant gets the element from fertile soil, put a check in the "Fertile soil" column.

You mean this female avocado pit will grow into a plant . . . in this glass of water?
That's hydroponics!
Is there any such thing as a male avocado pit?

PLANT NUTRIENTS

Essential Chemical Elements	SOURCES		
	Hydroponic compounds	Air	Fertile soil
Carbon, C			
Hydrogen, H			
Oxygen, O			
Nitrogen, N			
Phosphorus, P			
Sulfur, S			
Potassium, K			
Calcium, Ca			
Magnesium, Mg			
Iron, Fe			
Manganese, Mn			
Copper, Cu			
Zinc, Zn			
Boron, B			
Molybdenum, Mo			
Chlorine, Cl			
Iodine, I			

Questions to answer

1. Why do plants need rain or irrigation water for growth? ______

______

2. What chemical elements does the plant need to make starches, sugars, and fats? ______

What additional element to make protein? ______

______

3. Why are fertilizers added to soil? ______

______

4. What material on a farm is used as a "natural fertilizer?" ______

______

5. What are commercial fertilizers? ______

______

Extra for experts

1. Find out what a compost pile is and how it is made. Use reference books in the library or ask a friend who does a great deal of gardening.

Report to your class. ______

______

______

2. Compost is sometimes called "green fertilizer."

How is it used? ______

______

______

______

Tracer Elements in Plants

Radioactive tracer elements are used in plant research. An element used by the plant is made radioactive. The radioactive element is introduced into the plant. Its movements are traced by following the path of the radiation emitted. Thus a great deal can be learned about life processes in the plant.

Radioactive elements contain atoms which break up to emit radiation. There are three types of radiation:

(1) Alpha particles which are helium nuclei
(2) Beta particles which are electrons
(3) Gamma rays which are X rays of very short wavelength

Radiation from radioactive elements can be detected with a Geiger counter. Radioactive sources can be demonstrated in the laboratory. These give off beta particles and gamma rays. There is no danger to health since the amount of radiation is small.

One type of *nuclear demonstrator* is illustrated in Fig. 108. This includes an extendable probe with a Geiger counter at the end. The Geiger counter detects the radiation and the apparatus "reports" it by clicks and also by flashes of a small neon lamp. The number of particles which strike the Geiger counter may be estimated from the number of clicks and flashes. A rapid stream of clicks which seems to be a continuous sound would indicate that a large number of atoms are breaking up. Such a source of radiation would be very dangerous to the health of any one nearby.

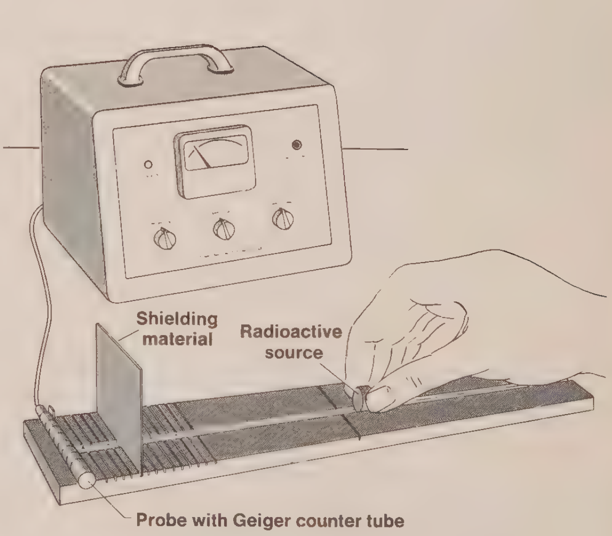
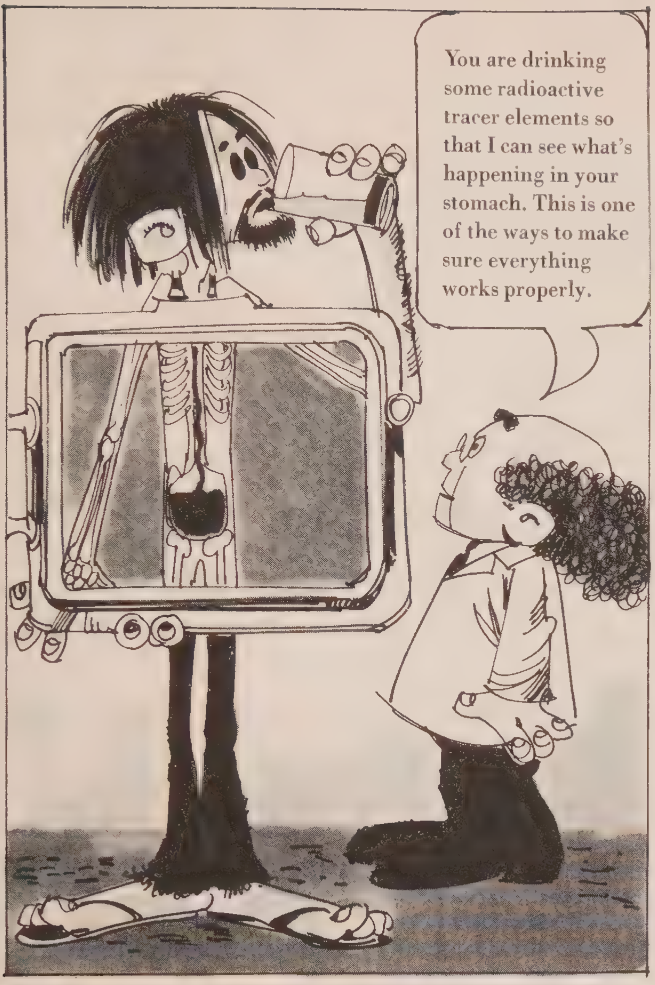


Fig. 108.

Experiment 21.6

Equipment and Supplies per Lab Station

Nuclear Demonstrator (Classmaster Kit)
radioactive sources
wristwatch with illuminated dial


You are drinking some radioactive tracer elements so that I can see what's happening in your stomach. This is one of the ways to make sure everything works properly.

What to do

1. Connect the nuclear demonstrator to a source of 110-volt alternating current.
2. Check for *background count.* Hold the Geiger counter probe in your hand at arm's length for a minute or two. Listen for clicks. These clicks are called BACKGROUND. Observe how often they occur at a particular time and place. Cosmic rays traveling through space will cause occasional clicks and flashes.

   Observations:

3. Detecting radioactivity
   (a) Place a radioactive source at one end of the grooved board. See Fig. 108. Listen for clicks and watch for flashes of light. Move the radioactive source nearer the probe a little at a time. Take a count of clicks at each position. How does the distance from the radioactive source influence the count made by the Geiger probe?

   Position (1)

   Position (2)

   Position (3)

   (b) Repeat the test using another radioactive source. Do some radioactive sources cause more clicks than others?

   Position (1)

   Position (2)

   Position (3)

   (c) Place a wristwatch with an illuminated dial at one end of the mounting fixture. Move it toward the Geiger counter just as you did the radioactive sources.

   Observations:

4. Enclosing the probe
   (a) Slip the cardboard cylinder around the Geiger probe. Repeat step 3, Parts (a), (b), and (c).
   Does the cardboard make any difference in the results?

   (1)

   (2)

   (3)

(b) Try the effect of the aluminum cylinder as in (a) above.

(1) ______________________________

(2) ______________________________

(3) ______________________________

(c) Try the effect of the lead cylinder as in (a) and (b) above.

(1) ______________________________

(2) ______________________________

(3) ______________________________

5. Thickness of shield

(a) Try different thicknesses of cardboard as a shield between the radioactive source and the probe. Place the source at the end of the mounting fixture.

Observations. ______________________

____

____

____

(b) Repeat (a) above first using squares of aluminum and then squares of lead. Is it possible to use the shields to prevent noticeable radiation reaching the Geiger probe?

Observations. ______________________

____

____

____

____

____

Questions to answer

1. How do you show that a material is radioactive? ______________________

____

2. How is the intensity of radiation at a particular spot affected by distance? ______________________

____

3. (a) Which of the three materials you tried seems to be the most effective as a shield for radiation? ______________________

(b) The next most effective? ______________________

(c) The least effective? ______________________

4. Suppose there had been a nuclear explosion in your part of the world. How could you check possible radioactive pollution of a certain field in the area? ______________________

____

____

RADIOACTIVITY

When a substance contains atoms which have nuclei which decompose giving off radiation, the substance is said to be RADIOACTIVE. In every case where a nucleus is unstable, it gives off one or more of the following kinds of radiation:

(1) Alpha particles which are helium nuclei, $^{4}_{2}He$. They consist of two protons and two neutrons.
(2) Beta particles which are electrons.
(3) Gamma rays which are a form of short wave radiation.

Alpha-decay: When a nucleus gives off an alpha particle, it loses two protons and two neutrons at the same time. Therefore it is converted into another element with an atomic number two units less and an atomic mass four units less.

The symbols used in writing nuclear equations show both the atomic number and the mass number with the symbol for the element. For example, $^{235}_{92}U$ represents uranium with a mass number of 235 and an atomic number of 92.

If $^{235}_{92}U$ undergoes alpha decay, it becomes an isotope with mass number 231 of an element with an atomic number of 90. The element with atomic number 90 is thorium. The nuclear equation may be written as follows: $^{235}_{92}U \rightarrow {}^{231}_{90}Th + {}^{4}_{2}He$

Beta-decay: In beta-decay, the nucleus gives off an electron. You may suppose that in some way a neutron separates into a proton which stays in the nucleus and the electron which is given off. Then, if a nucleus undergoes beta-decay, its atomic number increases by one unit but its mass number does not change.

Cobalt-60, $^{60}_{27}Co$, decomposes by beta-decay. When it does, the isotope of mass number 60 of the element with one higher atomic number, nickel, is formed. The nuclear equation may be written as follows:

$$^{60}_{27}Co \rightarrow {}^{60}_{28}Ni + e^{-}$$

Questions to answer

1. Explain each of the following symbols:

(a) $^{235}_{92}U$ ______________________________

(b) $^{231}_{90}Th$ ______________________________

(c) $^{4}_{2}He$ ______________________________

(d) $^{60}_{27}Co$ ______________________________

(e) $^{60}_{28}Ni$ ______________________________

(f) e ______________________________

2. How does alpha radiation differ from beta radiation in radioactive substances? __________

____

____

Laboratory Experiments in

ACTION CHEMISTRY

UNIT IX

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APPENDIX

Chapter 22

The Temperature of Flames

Fuels burn to give heat and light. The gas in your laboratory burner and a candle are fuels. Both are hydrocarbons. The gas in methane, CH_4. The candle is paraffin, $C_{25}H_{52}$, a white solid wax.

The chemical reaction of burning is COMBUSTION. In this experiment you will study the temperature of the flame of methane and of candle wax.

Use the color of copper at various temperatures as a guide to the temperature of the candle flame and the burner flame.

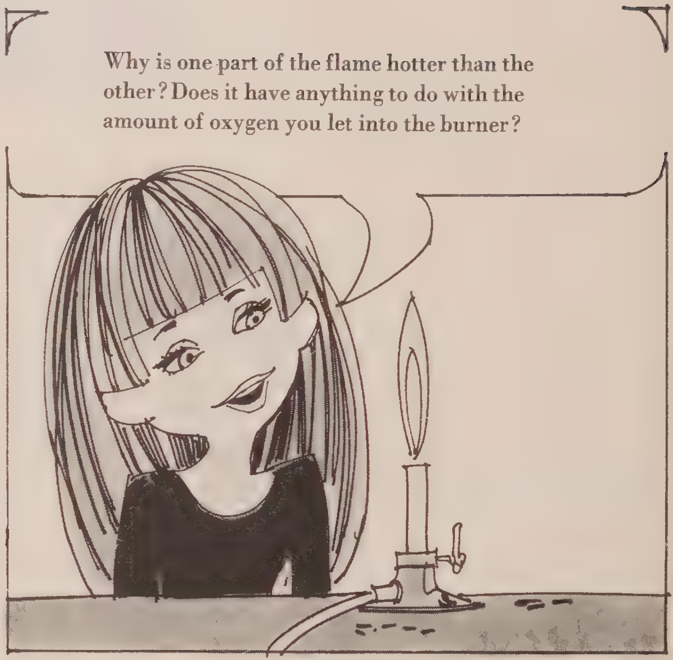
Color of Copper at Various Temperatures

Copper begins to show a faint red tinge at ---	500°C
Copper becomes dark red at ---------------------	650°C
Copper turns orange at ----------------------------	900°C
Copper melts at about -----------------------------	1100°C

Experiment 22.1

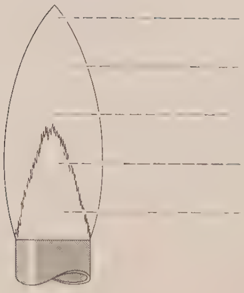
Equipment and Supplies per Lab Station

burner; glass square; test tube holder; tongs
candle, white paraffin, about 6 inches
copper wire, 16 or 18 gauge, 6 inches
cork—No. 6; sandpaper


Why is one part of the flame hotter than the other? Does it have anything to do with the amount of oxygen you let into the burner?

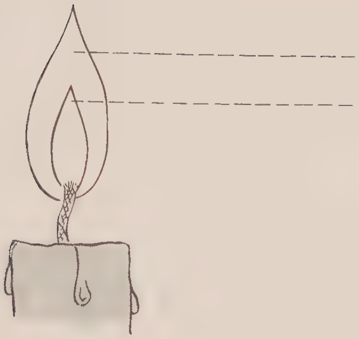
What to do

1. Obtain a six-inch length of heavy copper wire. Sandpaper it until it is bright and shiny. Make a holder by forcing one end of the wire securely into a cork.
2. Light the laboratory burner. Hold the tip of the wire in four or five of the positions shown in the illustration. Hold the wire in each position for two seconds. Compare the color of the copper with the table above. Record the color of the copper and the likely temperature in the position in the illustration.



Color of the Copper	Estimated Temperature
____________________	____________________°C
____________________	____________________°C
____________________	____________________°C
____________________	____________________°C
____________________	____________________°C

3. Light a candle. Attach it to a glass plate with a few drops of melted candle wax.



4. Repeat step 2 using the candle flame instead of the burner flame. Test two positions.

Color of the Copper	Estimated Temperature
______________	______________°C
______________	______________°C

Questions to answer

1. Which flame showed the higher temperature?

   ____

2. Which part of each flame showed the highest temperature?

   (a) burner flame ______________________

   (c) candle flame ______________________

3. Write balanced equations for combustion reactions you observed.

   (a) The burner flame: $CH_4 +$ ______________ →

   ______________ + ______________

   (b) The candle flame: $C_{25}H_{52} +$ ______________ →

   ______________ + ______________

4. Three fuels widely used in industry to produce heat energy are coal, fuel oil, and natural gas (methane).

   (a) Name one chemical element present in all three. ______________________

   (b) Name one product of combustion you would expect from all three. ______________

Extra for experts

Gasoline is a mixture of fuels used in automobiles. One of the compounds in the mixture is octane, C_8H_{18}.

1. What is the purpose of the carburetor in an automobile? ______________________

   ____

2. How is the gas mixture in the cylinder ignited?

   ____

3. Write the balanced equation for the combustion of octane.

   $C_8H_{18} +$ ______________ → ______________ +

   ____

4. Explain the power that pushes the piston and makes the automobile move. ______________

   ____

   ____

   ____

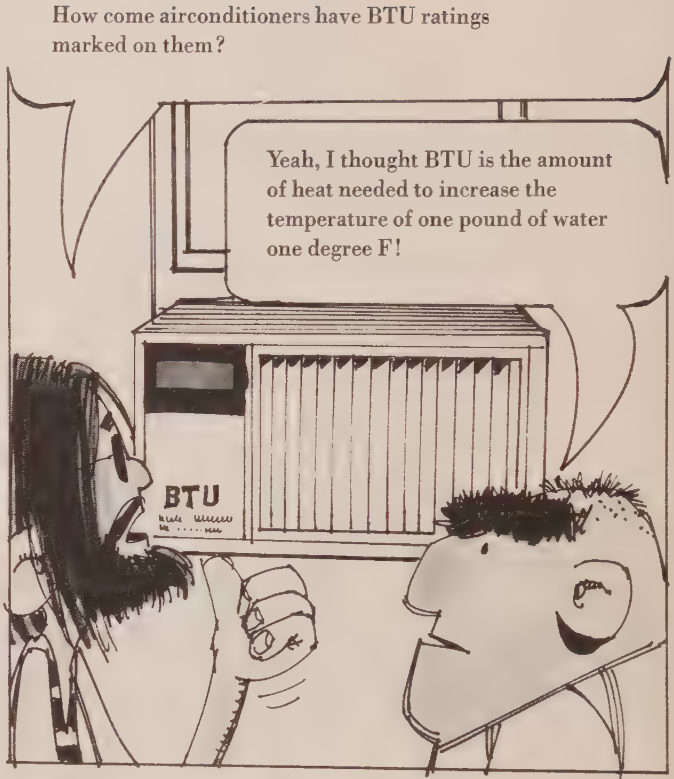
Fuels and Heat Energy

Fuels are rated according to the quantity of heat energy they can deliver. The heat production of industrial fuels is usually measured in British Thermal Units, BTU. In chemistry, heat energy is measured in calories as you learned in Chapter 16.

One BRITISH THERMAL UNIT, BTU, is the amount of heat needed to increase the temperature of one pound of pure water one Fahrenheit degree.

One CALORIE, cal, is the amount of heat needed to increase the temperature of one gram of pure water one Celsius degree. One British Thermal Unit, BTU, is about 252 calories.

"Canned heat" may be a very useful fuel on a camping trip or in any place where there is no gas or electricity. In this experiment you will make "canned heat" and determine the calories per mole it produces. You will then change these calories to British Thermal Units.


How come airconditioners have BTU ratings marked on them?
Yeah, I thought BTU is the amount of heat needed to increase the temperature of one pound of water one degree F!
BTU

Experiment 22.2

Equipment and Supplies per Lab Station

beaker—250 ml (if no small can is available)
Erlenmeyer flask—250 ml; glass rod; thermometer
buret clamp; ring stand
can (small vegetable can) about 3 inches high and 2½ inches in diameter
rubber band; aluminum foil, 3½ in x 3½ in
10 ml of a saturated solution of calcium acetate, $Ca(CH_3COO)_2$
60 ml of ethanol, denaturated; distilled water
ice cubes

What to do

PART I

1. Use a small, clean, dry vegetable can about 3 inches high and 2½ inches in diameter for a container.
2. Place in the can 10 ml of a saturated solution of calcium acetate, $Ca(CH_3COO)_2$. Add 60 ml of ethanol, CH_3CH_2OH. Stir with a glass rod.

   What happens? ____________________

   ____

   ____

   ____

   ____
3. Cover the can containing the product with a piece of aluminum foil, 3½ in. × 3½ in. Fold the aluminum down over the top of the can. Fasten securely with a rubber band.

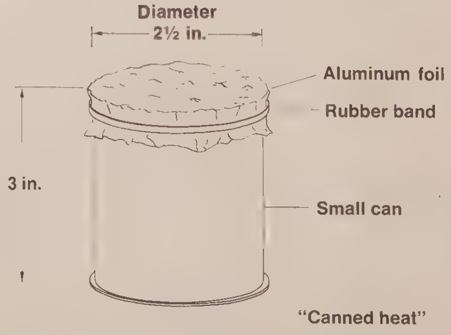


Fig. 109. Vegetable can containing "canned heat."

PART II

1. Set up the apparatus as illustrated in Fig. 110. Make sure the clamp on the Erlenmeyer flask is securely fastened.
2. Cool a little more than 200 ml of distilled water to between 10°C and 15°C. Use ice on the outside of the container only. Transfer exactly 200 ml of the cold water to the Erlenmeyer flask.
3. Remove the rubber band from the can of "canned heat". Weigh the can, cover, and "canned heat." Record the mass in the Data Chart.
4. Place the can under the flask of water. Make sure the flask is about 10 cm above the "canned heat." When the "canned heat" is lighted, the flask should be about 2 cm above the flame. *Note:* The flame itself is spread out and flickering, but may reach as much as 8 cm above the fuel. Then 8 cm + 2 cm = 10 cm from the surface of the fuel to the bottom of the flask.
5. Read the temperature of the water and record it in the Data Chart.
6. Remove the cover from the can. Light the "canned heat."
7. Burn the fuel until the temperature of the water is between 40°C and 50°C. Stir the water gently with the thermometer during the heating. Be sure that the bulb of the thermometer does not touch the glass of the flask.
8. Cover the can quickly and push it aside. Keep on stirring the water with the thermometer as long as the temperature continues to rise. When the temperature just starts to fall, record the highest temperature observed in the Data Chart.
9. Cool the can, cover, and remaining "canned heat" to room temperature. Weigh them and record the mass in the Data Chart.

DATA CHART

Mass of can, cover, and "canned heat" *before* burning	____g
Mass of can, cover, and "canned heat" *after* burning	____g
*Calculated mass of "canned heat" burned *(subtract the grams *after* burning from the grams *before*)	____g
Calculated mass of water heated (200 ml = 200 g) (Density of water = 1 g/1 ml, that is 1 ml weighs 1 g)	200 g
Temperature of the water at the *start* of the burning	____°C
Temperature of the water at the *end* of the burning	____°C
**Calculated* temperature change of the water *(subtract the °C at the *start* from the °C at the *end*)	____C°

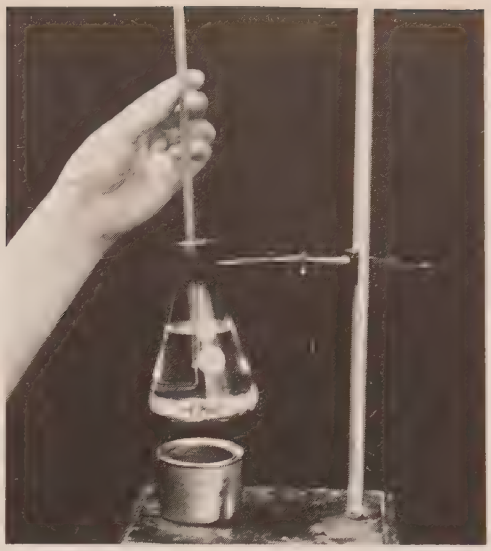


Fig. 110. Set up for PART II. The photo shows commercial "canned heat." (*Clara Aich and George Senty*)

Calculations

1. Calculated quantities from the Data Chart you will need
   - (a) Mass of "canned heat" burned ________g
   - (b) Mass of water heated 200 g
   - (c) Temperature change of the water ________C°
2. Calculation of calories (cal) produced by mass of "canned heat" burned

   cal = (200 g water heated) × (____________C° temp. change) × 1 cal/C°/1 g

   = ______________ calories produced by grams of "canned heat" burned
3. Calculation of calories (cal) per mole of the fuel in "canned heat"
   - (a) The fuel burned in "canned heat" is ethanol, CH_3CH_2OH

     g/mole of CH_3CH_2OH = (2 C × 12) + (6 H × 1) + (1 O × 16)

     = 46 g/mole of ethanol
   - (b) Calories produced per mole of ethanol burned

     $$\left(\_\_\_\_\_\_\text{cal produced}\right) \times \frac{46 \text{ g/mole of ethanol}}{\_\_\_\_\text{g ethanol burned}} = \_\_\_\_\_\_\text{cal/mole}$$
4. Calculation of British Thermal Units (BTU) per mole of ethanol
   - (a) About 252 calories = 1 BTU
   - (b) $$\frac{\_\_\_\_\_\_\text{cal/mole}}{252 \text{ cal/BTU}} =$$

     ____________BTU/mole of ethanol

Slippery Fluids

Out of oil miles from anywhere! What a disaster! Car engines simply will not run without oil.

Motor oils are lubricants. *LUBRICANTS are substances which make the surfaces of moving parts of machinery slippery.* They reduce friction. *FRICTION is the rubbing of one part against another.* Lubricated surfaces slide past each other rather than rub. Surfaces not properly lubricated may not only rub, but rub too hard. They may even stick together and quit moving. The car stops.

Motor oils, like honey, are viscous fluids. *Viscous* fluids do not flow readily. It takes some time to empty a can of motor oil into a crankcase. Viscous oils stick to metal surfaces. The surfaces stay slippery.

Does the viscosity of motor oils vary with the "weight?"

Does heat make a difference in the viscosity of a motor oil?

Test some motor oils and find out.

Experiment 22.3

Equipment and Supplies per Lab Station

(3) beakers – 250 ml
graduated cylinder – 25 ml
(2) funnel assemblies *each one* consisting of:
small, short-stemmed funnel, 3 in. length of rubber tubing, glass tip from medicine droppers, pinch clamp
burner; acetone or denatured alcohol
ring stand with (2) small rings and (2) clay triangles
ring stand with small ring and wire gauze
test tube – 18 mm × 150 mm; distilled water
30 ml motor oil (one of three different samples assigned to each pair of students)

What to do

First day: Prepare the equipment to be used in this experiment at least one day before it is used.

1. Connect a small, short-stemmed funnel, a three-inch piece of rubber tubing, and the glass tip from a medicine dropper. Place a pinch clamp on the rubber tubing. See Fig. 111.

   Make sure that all parts of the assembly are clean.
2. Place a mark at the point where the stem of the funnel begins. Use a felt pen or a narrow strip of masking tape. Label this No. 1.

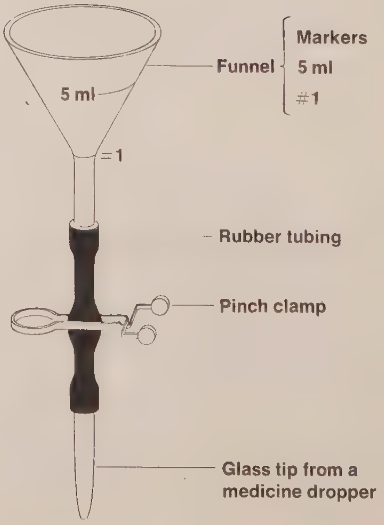


Fig. 111.

3. Place the funnel in a support with a beaker under the delivery tip. See Fig. 112. Fill the funnel with water.
4. Open the pinch clamp to allow the water to fill the stem, rubber tubing, and delivery tip. Force air bubbles out by raising the delivery tip with the clamp open slightly.
5. Lower the water level down to the No. 1 mark.
6. Transfer 5 ml of water to the funnel. Place a third marker at this level. Label it 5 ml.
7. Drain all of the water out of the funnel and connections. Remove the pinch clamp. Rinse the apparatus with acetone or denatured alcohol. Shake out as much liquid as possible.
8. Put the apparatus in your locker to dry thoroughly before your next laboratory period.

Second day: Two students work together. Each supplies a funnel assembly.

Testing Viscosity

9. Replace the pinch clamp on the rubber tubing. Place two funnel assemblies in funnel supports with a beaker under each.
10. Fill one funnel with water. Use the pinch clamp in filling the stem, rubber tubing, and delivery tip. Force air bubbles out.
11. Lower the level of the water to the 5 ml mark.
12. Follow steps 10 and 11 in filling the second funnel assembly to the 5 ml mark with the motor oil assigned to you by your teacher.
13. (a) *One partner*: Watch the second hand on a clock or a watch. Record the exact number of seconds it takes 5 ml of water to run out of the funnel.

    (b) *Other partner*: Open the pinch clamp wide. Drop the water level from 5 ml to the No. 1 mark. Close the pinch clamp instantly. This delivers just 5 ml of water.
14. Make three trials. Take the average number of seconds as your result. Record in Data Chart 1.

DATA CHART 1

First trial ____________seconds

Second trial ____________seconds

Third trial ____________seconds

Average ____________seconds for the water

Note:

$$\text{Average} = \frac{1\text{st} + 2\text{nd} + 3\text{rd}}{3}$$

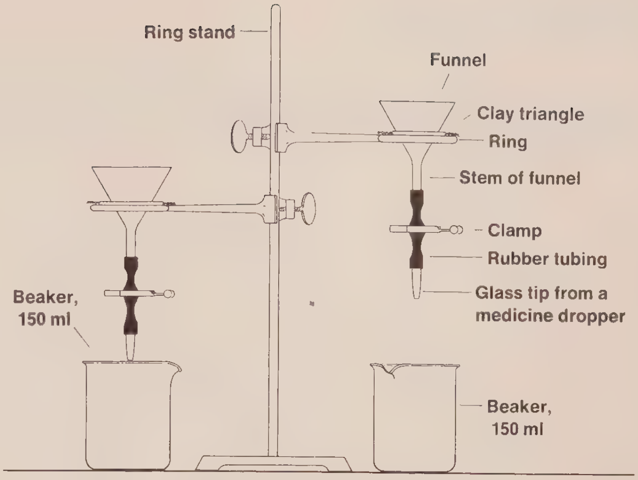


Fig. 112. Viscosity testing set up.

15. Repeat steps 13 and 14 in using the second funnel assembly. Determine the average number of seconds for the motor oil. Record in Data Chart 2.

DATA CHART 2

First trial ______________seconds

Second trial ______________seconds

Third trial ______________seconds

Average ______________seconds for the motor oil

16. Calculate the relative viscosity of the motor oil. *Relative viscosity* means how much more slowly the motor oil flows than the water.

$$\text{relative viscosity} = \frac{\text{ave. sec. for the motor oil}}{\text{ave. sec. for the water}}$$

$$= \underline{\hspace{6cm}}$$

Record the relative viscosity of your sample of motor oil in Data Chart 4. Obtain relative viscosities from students who have tested other motor oils and record.
17. Place 20 ml of your motor oil in a test tube. Stopper the test tube loosely. Heat the test tube in a beaker of boiling water for 10 minutes. CAUTION: DO NOT HEAT OIL OVER AN OPEN FLAME.
Note: If electric hot plates are available, place 20 ml of the motor oil in a beaker on a hot plate. Heat as directed by your teacher.
18. Drain all cold oil out of the second funnel assembly.

19. Repeat the procedures with the warm motor oil. Keep the oil not being tested warm between trials. Record results in DATA CHART 3.

DATA CHART 3

First trial ________ seconds

Second trial ________ seconds

Third trial ________ seconds

Average ________ seconds

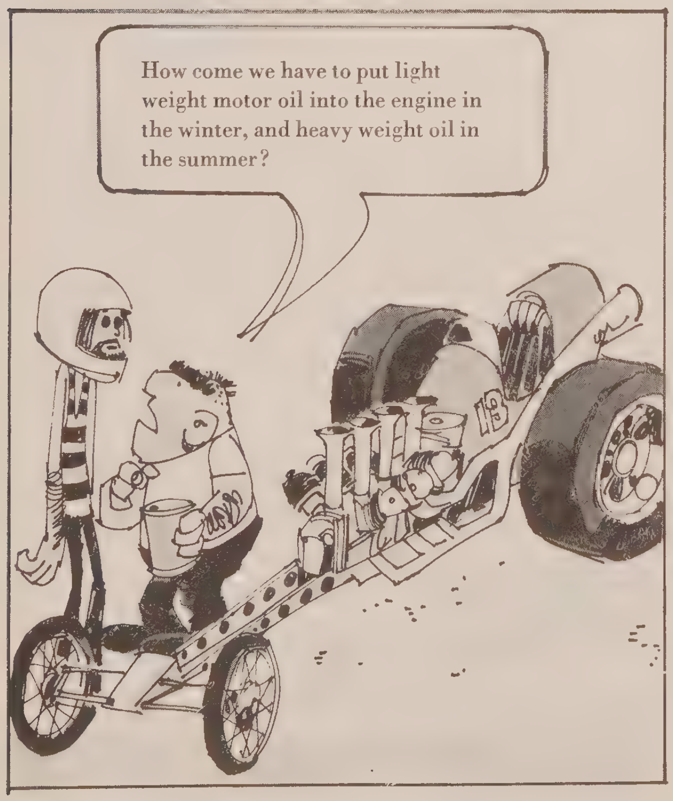
20. Calculate the relative viscosity

$$\text{relative viscosity} = \frac{\text{ave. sec. for warm motor oil}}{\text{ave. sec. for water}}$$

$$= \_\_\_\_\_\_\_\_$$

DATA CHART 4

	"Weight" Motor Oil	Average Time seconds	Relative Viscosity
Water			1.0
Motor Oil 1			
Motor Oil 2			
Motor Oil 3			


How come we have to put light weight motor oil into the engine in the winter, and heavy weight oil in the summer?
13

Questions to answer

1. How were the relative viscosities related to the "weight" by which the motor oil was sold? ________

____

2. What difference in relative viscosity did warming each of the motor oils make? ________

____

3. Why might a different "weight" of motor oil be be used during a hot summer and during a very cold winter? ________

____

Chapter 23

NATURAL RUBBER AND SYNTHETIC RUBBER

To obtain natural rubber, incisions are made in the bark of rubber trees. A milky liquid known as LATEX oozes out. In the manufacture of rubber, other ingredients are added to the latex.

Synthetic rubber is made from chemicals obtained largely from coal tar and petroleum. Neoprene is a common synthetic rubber used in the chemistry laboratory for stoppers and tubing.

In this experiment, compare some of the properties of natural rubber and synthetic rubber.

Experiment 23.1

Equipment and Supplies per Lab Station

beaker—250 ml; graduated cylinder—25 ml
(12) test tubes; burner; (2) test tube racks
ring stand with ring and wire gauze
6 ml benzene; 6 ml ethanol; 6 ml kerosene
6 ml hydrochloric acid, dilute
6 ml sulfuric acid, dilute
6 ml sodium hydroxide solution, dilute
samples of natural rubber and
synthetic rubber

What to do

1. Obtain samples of natural rubber and synthetic rubber. Test and compare the following properties.
   - (a) *Elasticity*: Do the two kinds of rubber have the same amount of stretch? ________
   - (b) *Squeeze*: Can each be squeezed as easily as the other? ________
   - (c) *Flexibility*: Does one bend and return to shape more easily than the other? ________
   - (d) Do both erase pencil marks? ________
   - (e) Drop a sample of each in boiling water for five minutes. How has each been affected by the heat? ________
   - (f) Examine a sample of each which has been in the refrigerator overnight. How has each been affected by the cold? ________

Why do they change to 'wet road tires' when it rains? Are the tires different?
5

2. Drop a small piece of natural rubber in each of six test tubes. Add 2 to 3 ml of each of the following reagents to different test tubes. Watch the tests for 15 minutes. (Start step 3 right away so that you can watch all tubes during the same 15 minutes.) Record your observations in the Data Chart.

(a) Ethanol	(d) Dilute sodium hydroxide, NaOH, solution
(b) Benzene	(e) Dilute hydrochloric acid, HCl, solution
(c) Kerosene	(f) Dilute sulfuric acid, H_2SO_4, solution

3. Repeat step 2, using synthetic rubber instead of natural rubber.

DATA CHART

Test Liquid	Natural Rubber	Synthetic Rubber
Ethanol		
Benzene		
Kerosene		
Dilute NaOH		
Dilute HCl		
Dilute H_2SO_4		

Questions to answer

1. Would natural or synthetic rubber be better for rubber bands, bathing caps, and rubber gloves?

Explain. ______________________________

____

2. Which of the two kinds of rubber would be better for stoppers for laboratory reagent bottles? Explain. ______________________________

____

3. Which of the two kinds of rubber would be better for soles and heels, overshoes, and raincoats? Explain. ______________________________

____

4. Would you line the hose of a gasoline pump with natural rubber or neoprene? Explain. (*Hint*: Gasoline may act like benzene or kerosene.) ______________________________

____

MAKING A PLASTIC SIMILAR TO BAKELITE

Bakelite is used for radio cabinets, electrical insulators, varnishes and enamels, and even for plastic jewelry.

A plastic similar to bakelite can be made in the laboratory from aniline hydrochloride and forty percent formalin solution.

Demonstration 23-A

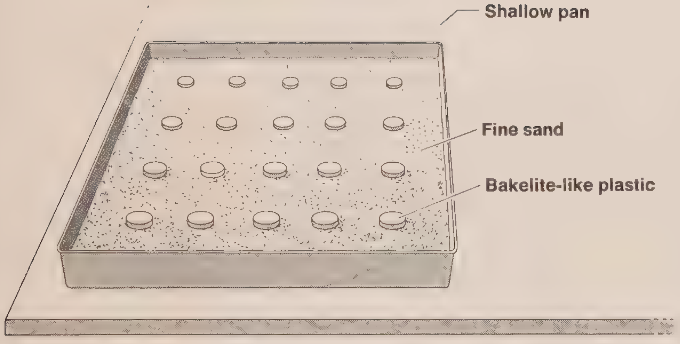


Fig. 113. Allow plastic material to harden in mold.

Equipment and Supplies per Demonstration

container-wide mouthed – 300 ml; glass stirring rod
graduated cylinder – 25 ml; (2) graduated cylinders – 100 ml
balance, laboratory; hammer; piece of marble with sharp edges; samples of articles made from bakelite
sand mold, triangular file,
saturated solution of aniline hydrochloride, 100 ml
40% formalin solution, 100 ml

What to do

1. Pour at exactly the same instant into a wide-mouthed container (*Hint*: Hold the container of one reagent in each hand. These solutions should be prepared beforehand.)
   100 ml of a saturated solution of aniline hydrochloride
   100 ml of 40% formalin solution
2. Stir well with a glass rod.
3. Pour small amounts of the viscous material into each of several small depressions in a sand mold. Allow it to harden.
4. Students examine the product.
   (a) What color is it? ______________________
   (b) How hard is it? Try scratching it with marble. ______________________
   ____
   (c) What is its density? ______________________
   (d) Try sawing it with a triangular file. What happens? ______________________
   ____
   (e) Hit a small piece with a hammer. Is it brittle or malleable? ______________________

Question to answer

List as many common useful things as you can that are made of bakelite or with parts made of bakelite.

____

____

____

Saran Wrap and Sandwich Bags

Saran Wrap and sandwich bags are both used to keep foods fresh. They look very much alike, but they differ in the chemicals from which they are made. Saran Wrap is a vinyl chloride polymer. Most sandwich bags are polyethylene.

How can you tell vinyl chloride polymer from polyethylene? Try burning a small sample of each. The smoke from vinyl chloride polymers contains hydrogen chloride. Hydrogen chloride dissolves in water to form hydrochloric acid. The acid lowers the pH of the water. Smoke from nonchloride polymers such as polyethylene does not lower the pH of the water as much.

Experiment 23.2

Equipment and Supplies per Lab Station

beaker—250 ml; funnel; burner; can; $1\frac{1}{2}$ in deep × $2\frac{1}{2}$ in diameter
clamp; ring stand; 8 in rubber tubing
1 g of polyethylene sandwich bags; 1 g of Saran Wrap
distilled water; Universal Indicator paper

What to do

1. Obtain a small shallow tin can like a deviled ham can. Make sure that the can is clean and dry.
2. Tie a strip of Saran Wrap into a knot. Weigh as close to one gram as possible. Place the weighed Saran Wrap in the can.
3. Mount an inverted funnel on a ring stand. Connect a piece of rubber tubing from the stem of the funnel into a beaker containing 100 ml of distilled water. See Fig. 114.
4. Place the apparatus under strong ventilation, as under a *hood*.
5. Use the flame of a burner to ignite the Saran. Allow the Saran to burn completely. The smoke will go into the water.
6. Test the pH of the water with Universal Indicator paper. Record the result in the Data Chart.
7. Clean the can and rinse the funnel and tubing. Place a fresh 100 ml of distilled water in the beaker.
8. Repeat steps 2–6 using one gram of polyethylene sandwich bag material instead of the Saran Wrap.
9. Repeat the experiment with as many samples of plastic food wraps, various containers, broken phonograph records, and similar materials as possible. Record the results.

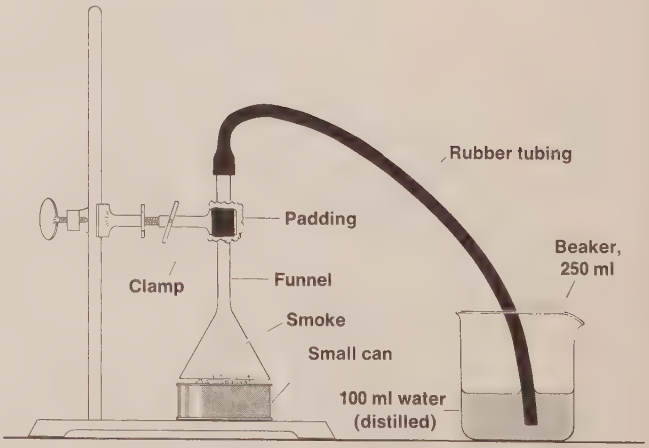
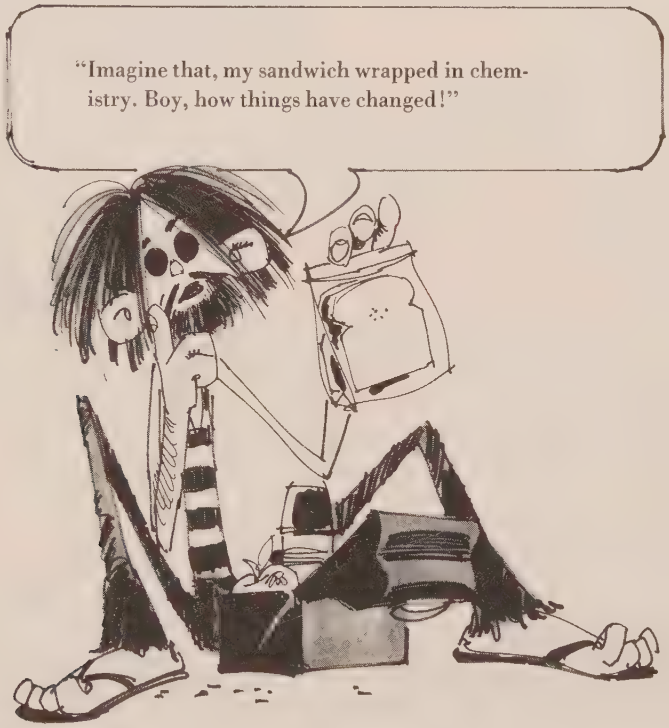


Fig. 114.

DATA CHART

Sample	pH of water	Possible compound
1. Saran Wrap		vinyl chloride polymer
2. Sandwich bag		polyethylene
3.		
4.		
5.		


"Imagine that, my sandwich wrapped in chemistry. Boy, how things have changed!"

Questions to answer

1. Find as many uses as you can for vinyl chloride polymers. ______

2. Find as many uses of polyethylene as you can. ______

THERMOPLASTICS AND COLD SETTING PLASTICS

Thermoplastic materials can be softened by heat and molded as desired while warm. When cool they will harden in the molded shape. Sealing wax is a natural resin plastic.

Cold setting plastic once hardened is not softened by heat. "Bioplastic" is a cold setting plastic.

Test the properties of sealing wax and "Bioplastic" for yourself.

Experiment 23.3

Equipment and Supplies per Lab Station

beaker—250 ml; burner, laboratory
mold in which to "set" the "bioplastic"
ring stand with ring and wire gauze
materials for forming "bioplastic"
small object for mounting in the "bioplastic"
short piece from a stick of sealing wax

What to do

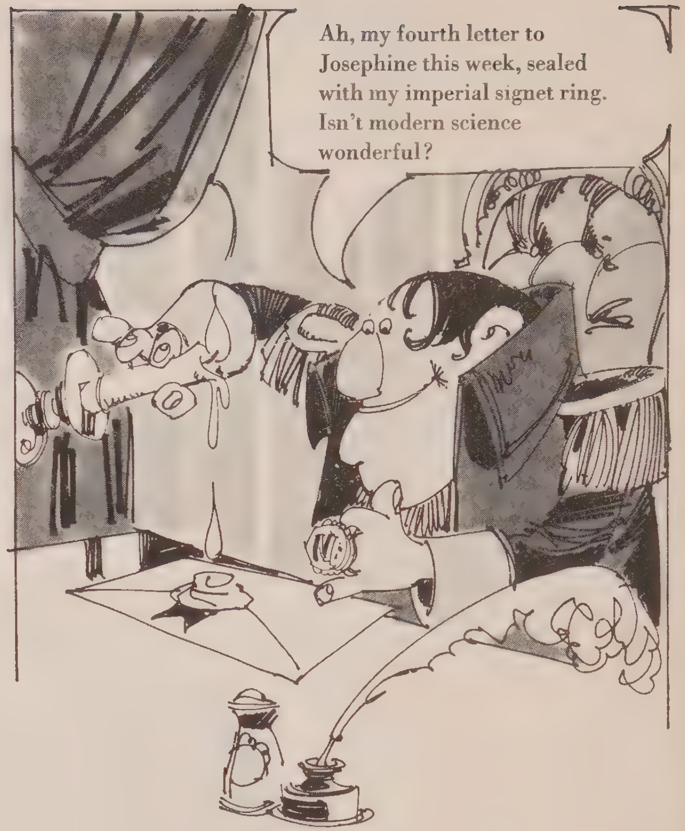
PART I

1. Warm some sealing wax. Let a few drops fall on paper. Imprint the soft wax with your class ring or some hard object which will leave a design.
2. Describe what happened. ______________________

____

PART II

1. Prepare cold setting "bioplastic" according to the manufacturer's directions. Set some small object as a flower, a leaf, or a blade of grass in the plastic before it hardens. Allow the plastic to harden.


Ah, my fourth letter to Josephine this week, sealed with my imperial signet ring. Isn't modern science wonderful?

2. Warm the plastic in boiling water. Try changing its shape. What happens? ______________

____

PART III

1. Select three to four plastic objects or materials to test and classify.
2. Immerse the object or material in boiling water for five or more minutes. The objects that can be softened by heat and bent or molded while warm are thermoplastic.

CLASSIFICATION

Sample	Thermoplastic	
Name	Yes	No: Explain
1.		
2.		
3.		
4.		

Questions to answer

1. Thermoplastics are suitable for what uses, in general? ______________________

____

2. Cold setting plastics are used for what kinds of articles? ______________________

____

More About Plastics

Test the properties of a number of different plastic objects or materials found around the home.

Experiment 23.4

Equipment and Supplie's per Lab Station
(6) test tubes (if washed and reused, otherwise, 36)
burner; forceps; hammer or similar heavy object;
test tube rack
large assortment of plastic objects and materials used around homes
distilled water; 30 ml ethanol; 30 ml dilute hydrochloric acid
30 ml dilute sulfuric acid; 30 ml dilute sodium hydroxide

What to do

1. Obtain three or more samples from plastic objects or materials used in homes.
2. Record the following properties for each in Data Chart 1.
   (a) Color
   (b) Transparency—can you see through it?
   (c) Flammability—does a very small sample held by a forceps in the burner flame burn easily?
   (d) Hardness—does it scratch easily?
   (e) Elasticity—does it stretch and return to shape?
   (f) Flexibility—does it bend without breaking?
   (g) Brittle or malleable—does it crumble when hit by a heavy object?

DATA CHART 1

Sample	Color	Transparent	Flammable	Hardness	Elastic	Flexible	Brittle or malleable
1.							
2.							
3.							
4.							
5.							

3. Test the solubility of about one gram of each sample in 5 ml of each of these reagents: water, ethanol, dilute sodium hydroxide, (NaOH), dilute hydrochloric acid (HCl), and dilute sulfuric acid (H_2SO_4). Record your observations in Data Chart 2.
   Note: Place about one gram of *each* plastic sample in *each* of *five* test tubes. Add one of the reagents to each test tube. Shake well and observe.

DATA CHART 2

Sample	Water	Ethanol	NaOH	HCl	H_2SO_4
1.					
2.					
3.					
4.					
5.					

Plastic bags are good because they make collecting garbage easy!

Questions to answer

1. Must care be used in bringing some plastic objects near direct heat? Explain. ______
2. How durable are the plastic objects you have observed compared with similar objects, toys for example, made of metal or of wood? ______
3. What are some advantages of plastics that make them used so much? ______
4. What plastic materials, if any, could be used as containers for storage of the following chemicals? (Consider resistance to breaking as well as solubility.)
   - (a) ethanol ______
   - (b) bases, such as sodium hydroxide ______
   - (c) hydrochloric acid ______
   - (d) sulfuric acid ______
5. Write a paragraph relating the properties of plastics to their uses. ______

Chapter 24

Natural Fibers and Synthetic Fibers

Fabrics are woven from fibers, some natural and some synthetic. Think of the infinite variety of fabrics used in clothes, in homes, and in industry!

Natural fibers include cotton from cotton plants, linen from flax plants, wool from sheep and goats, and silk from silkworms.

Synthetic fibers are products of the chemical industry. Familiar ones are acetate silk, polyester, nylon, orlon, dacron, lycra, and spandex.

How are the many fibers alike and how are they different? Make a few tests and see for yourself.

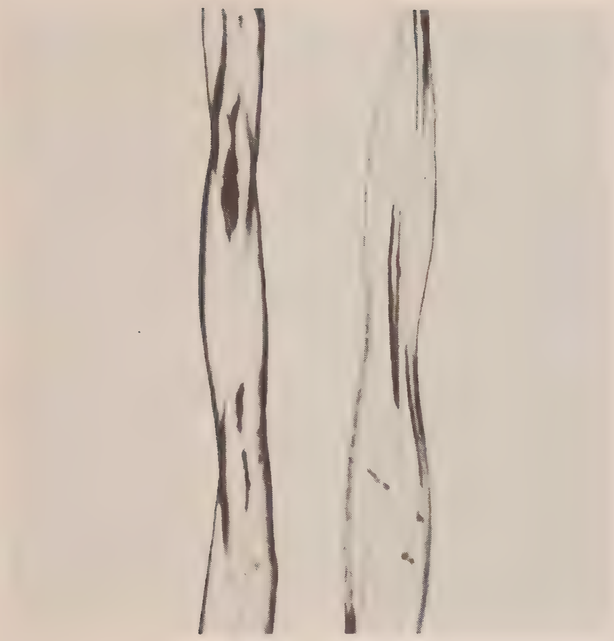


Fig. 115. Photomicrograph of cotton fibers. (*Courtesy of National Cotton Council*)

To clothe a population of 200 million, we have to produce 400 million square yards of fabric a year. And that's just enough material for one suit per person. Thank goodness for synthetic fibers. Without it we couldn't clothe everybody!

Experiment 24.1

Equipment and Supplies per Lab Station

burner, laboratory; forceps; medicine dropper; microscope or hand magnifier
samples of pure cotton, linen, wool, and silk
samples of two or more synthetic fabrics

What to do

1. Obtain samples of cotton, linen, wool, silk, and two or more synthetic fabrics.
2. Fray out a few fibers from each sample and examine them under the low power of a microscope or with a hand magnifier. Make a sketch and describe each. Ask your teacher for an unknown fabric. Decide from what kind of fibers it is woven.

(a) Cotton

(b) Linen

(c) Wool

(d) Silk

(e)

(f)

(g) Unknown

What fabric is it? ________________

3. *Tensile strength:* Pull hard on one of each kind of fiber, first when it is dry and then when wet. How strong is it, both dry and wet? Record in the Data Chart "++++" for the one hardest to break, "+" for the one easiest to break, and "+++" and "++" for those in between.
4. *Elasticity:* Test the ability of each kind of fiber to stretch and return to shape. Record in the Data Chart as '+'s" as in Step 3. The one with the greatest elasticity is "++++".
5. *Combustibility:* Gather in forceps a few fibers from each sample in turn. Hold each set of fibers in a flame. Record in the Data Chart what happens.
6. *Water absorption:* Sprinkle a drop or two of water on a piece of each fabric. Record in the Data Chart whether the water soaks into the fabric or tends to run off.

Which one dries the fastest? ____________________

Fig. 116. Photomicrograph of wool fibers.

DATA CHART

Sample	Tensile strength		Elasticity	Combustibility	Water absorption
	Dry	Wet			
Cotton					
Linen					
Wool					
Silk					
?					
?					
Unknown					

Questions to answer

1. Look at the fibers and fabrics carefully.
   (a) Why do you think winter coats are made of wool? ____________________
   ____
   (b) Why is cotton cool to wear in summer?
   ____
   ____
2. Which fabric would be best for making parachutes? Explain. ____________________
   ____
3. Which fabric, or fabrics, would make the best fitting bathing suits? ____________________
4. Which fabrics would be safest for making children's pajamas? (Consider the fire hazard.)
   ____
5. Which fabric is best for quick-drying clothes?
   ____

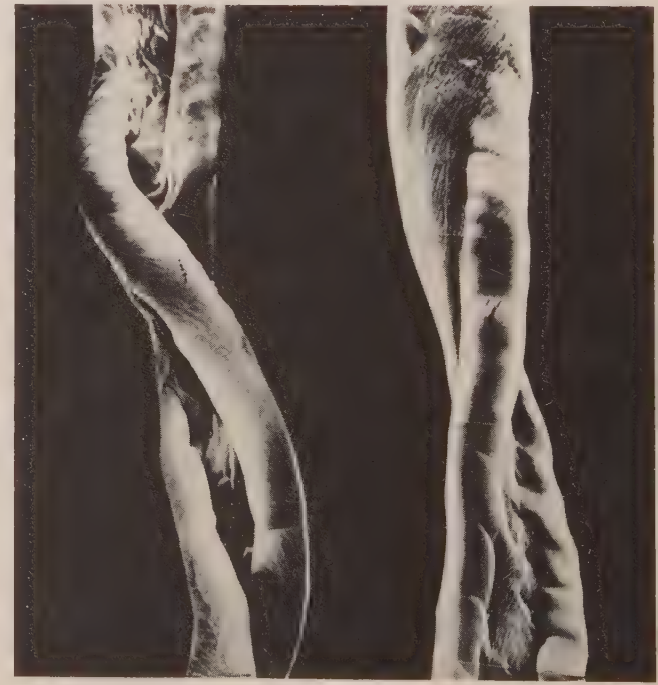


Fig. 117. Photomicrograph of nylon fibers. (*Peter Arnold, Agent*)

What Are Fibers Made Of?

Cotton and linen fibers are from plants; wool and silk fibers from animals. In which of these fibers would you expect to find starches and sugars? In which ones would you expect to find proteins? In the experiments in Chapter 16, what chemical elements did you find to be present in starches and sugars and which ones in proteins?

Synthetic fibers are made from chemicals. Would you care to make a guess about the chemical elements in them?

This experiment may help you find some answers to these questions.

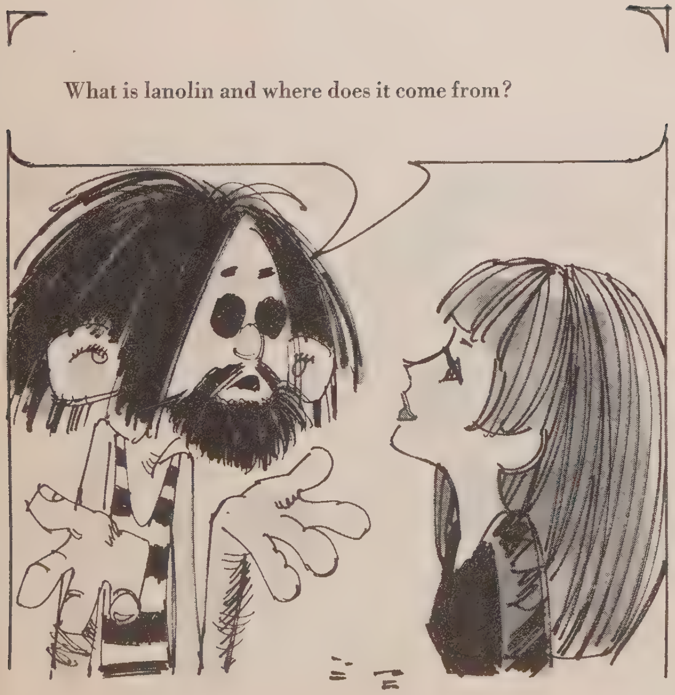
Experiment 24.2

Equipment and Supplies per Lab Station

burner, laboratory; clamp, buret; ring stand
(6) test tubes, 18 mm × 150 mm
small dropper bottle of lead acetate solution
(6) strips each of red and blue litmus paper
samples of fabrics, cotton, linen, wool, silk, synthetics
copper wire, fine, 22 gauge, $1\frac{1}{2}$ inches; filter paper

What to do

1. Place a small piece of each fabric to be tested in a separate test tube. Label each test tube. Test samples of cotton, linen, wool, silk, and synthetics.
2. Arrange each test tube in turn as you did in Experiment 16.5, Fig. 81 B, page *6*/**27**.
   - (a) Suspend a strip of damp red and a strip of damp blue litmus paper in the top of the test tube so that the two pieces of paper cannot touch. (What color change with the litmus shows the presence of nitrogen? ____)
   - (b) Cover the test tube with filter paper dampened with lead acetate solution. (What change in the lead acetate shows the presence of sulfur? ____________________)
3. Heat each test tube in turn. Record in the Data Chart, any moisture formed, any change in either litmus strip, any change in the lead acetate. (*Note:* Look for a change in **either** of the litmus papers.)
4. Examine the residue left in the test tube after the heating. Record its appearance in the Data Chart.


What is lanolin and where does it come from?

DATA CHART

Sample	Moisture formed	Litmus Reaction	Lead acetate	Residue-appearance
Cotton				
Linen				
Wool				
Silk				
?				
?				

Questions to answer

1. (a) With which samples did moisture form? ____________________

____

(b) What chemical elements must have been present? ____________________

____

2. Which samples gave an acid reaction with litmus? ____________________

____

Note: This reaction is found with cotton and linen.

3. Which samples gave both a basic reaction with litmus and a darkening of the lead acetate?

____

____

(a) These reactions show the presence of what two chemical elements? ____________________

(b) The presence of these two chemical elements shows the probable presence of what chemical compound? ____________________

(c) In which of the fabrics would you expect to find this chemical compound? Explain. ____________________

____

4. Is there any one chemical element which seems to be present in all of the fibers? ____________________

If so, what is it? ____________ Explain. ________

____

5. The synthetic fibers are most like which of the natural fibers? Explain. ____________________

____

6. Look up the materials from which each of these two synthetic fibers is made:

(a) Nylon ____________________

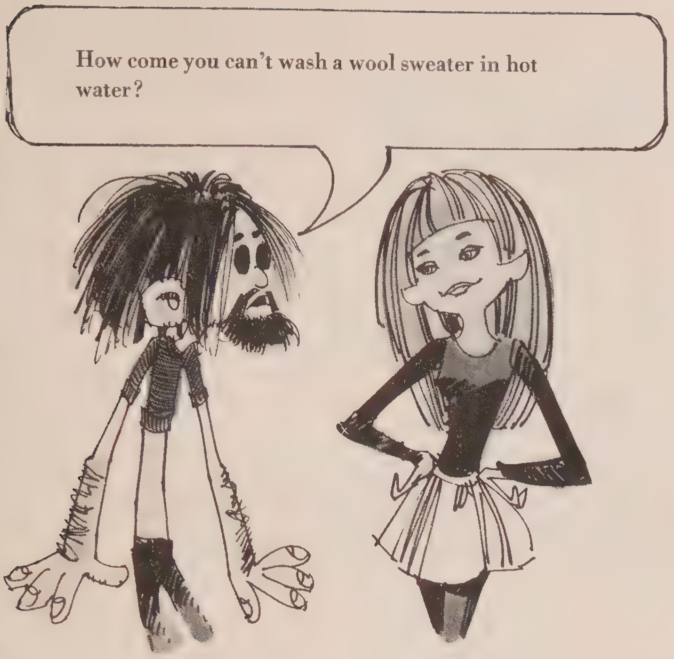
____

(b) Orlon ____________________

____

(c) Any other synthetic fiber in which you are interested. ____________________

____


How come you can't wash a wool sweater in hot water?

HOW FABRICS GET ALONG WITH CHEMICALS

Is it a good idea to follow the manufacturer's instructions about the care of a new sweater or dress or shirt?

Are laboratory precautions against acid and base damage to clothes as well as skin also sensible?

In this experiment try the effect of chemicals on fabrics.

Experiment 24.3

Equipment and Supplies per Lab Station

graduated cylinder—25 ml; (12) test tubes, 18 mm × 150 ml
forceps
stopper to fit the test tubes; test tube rack
60 ml of each of the following solutions: 3 *M* hydrochloric acid; 2 *M* sulfuric acid; 1 *M* sodium hydroxide; household bleach; soap solution; synthetic detergent
samples of cotton, linen, wool, silk, and of two or more synthetic fabrics

What to do

1. Test samples of cotton, linen, wool, silk, and two or more synthetic fabrics.
   Note: Your teacher may assign you particular fabrics to test or particular solutions to use in testing all fabrics. If so, save the samples you have tested for others to see. Be sure to observe and record results with all fabrics and all solutions.
2. Place a small piece of each fabric in a test tube. Add 10 ml of soap solution to each test tube. Shake vigorously. Let stand for five to ten minutes. Pour off the solution. Spread the samples on paper towel to dry. Record any changes observed. Note changes in color, appearance, tensile strength, and elasticity.
3. Repeat step 2. Add synthetic detergent solution instead of soap solution. *Hint to save time:*
   (a) Start this second set of tests while the first set is standing.
   (b) Observe and record any changes in step 2 set while step 3 is standing. Rinse out the test tubes so that they will be ready for step 4.
   (c) Continue this procedure until all fabrics have been tested with all solutions.
4. Repeat step 2. Add chlorine bleach instead of soap solution.
5. Repeat step 2. Add 1 *M* sodium hydroxide solution instead of soap solution.
6. Repeat step 2 with 3 *M* hydrochloric acid.
7. Repeat step 2 with 2 *M* sulfuric acid.

DATA CHART

Sample	Soap soln.	Detergent soln.	Bleach soln.	NaOH soln.	HCl soln.	H_2SO_4 soln.
Cotton						
Linen						
Wool						
Silk						
?						
?						

Questions to answer

1. (a) Did you find any fabrics which it would be safer to wash with soap than with synthetic detergent? Explain. ______________________________
   ____
   ____
   (b) Any safer to wash with synthetic detergent? __________ If so, which ones? ______________
   ____
   (c) Any with which neither soap nor detergent is safe? If so, which ones? ______________
   ____
2. Are there any of the fabrics on which you cannot use chlorine bleach safely? Explain. ____
   ____
   ____
3. Of the sodium hydroxide, the hydrochloric acid, and the sulfuric acid, which one seemed to be
   (a) the most damaging to the most fabrics? ____
   ____
   ____
   (b) the least damaging? ______________________
   ____

HOW ABOUT A LITTLE COLOR?

The clothes and other fabrics you live with every day would be pretty uninteresting without color. People have dyed cloth from the earliest known times.

Two general methods of dyeing cloth are commonly used.

Direct dyeing: The particles of the dye are affixed directly to the cloth. This is usually a less expensive method.

Indirect dyeing: First, a chemical substance known as a *mordant* is affixed to the cloth. The dye unites with the mordant. Mordants uses include: (1) Tartar emetic, potassium antimonyl tartrate, $KSbOC_4H_4O$, which forms an insoluble gelatinous material which fastens to the cloth. The dye unites with the mordant. (2) Mordants for basic dyes: Aluminum sulfate, $Al_2(SO_4)_3$ and salts of iron, chromium, and tin.

Experiment 24.4

Equipment and Supplies per Lab Station

(2) beakers—400 ml; graduated cylinder—100 ml
ring stand with ring and wire gauze
burner, laboratory; forceps
200 ml of each of the following solutions:
aluminum sulfate solution; ammonium hydroxide solution, 0.5 *M*; dye solution (Malachite green, fuchsin, or methyl orange); sodium carbonate solution, 0.25 *M*
(4) pieces of white cotton cloth, each about 2 inches square

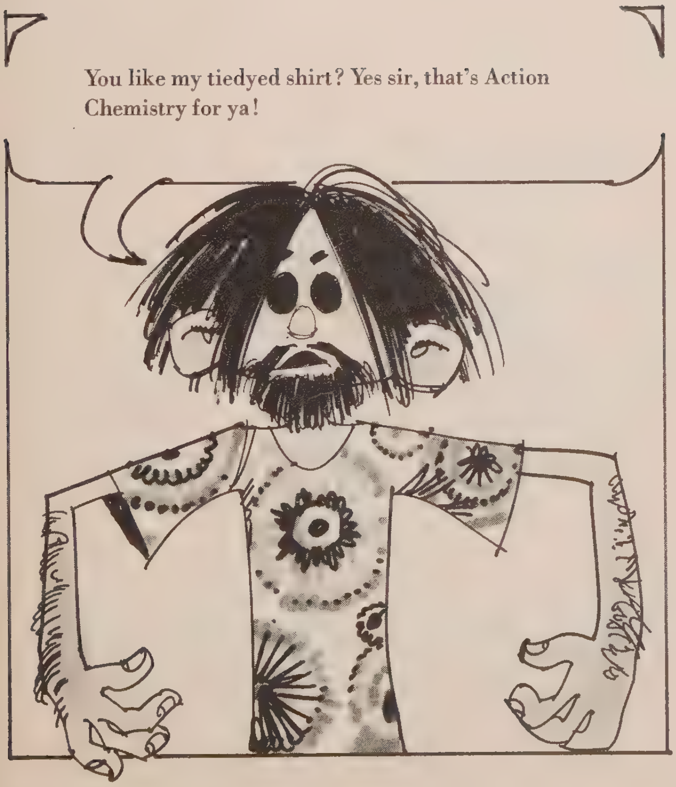
What to do

PART I REMOVING THE SIZING FROM CLOTH

1. Obtain four samples of white cotton cloth.
2. Place 200 ml of 0.25 *M* sodium carbonate solution in a 400-ml beaker. Put all four samples of the cloth into the solution. Bring to a boil. Boil for two minutes.
3. Decant the carbonate solution. Rinse the cloth samples in water. Wring them out. This process removes the sizing. *SIZING is material used to fill the pores of fibers of fabrics.*
4. Scrub and rinse the 400-ml beaker thoroughly so that it can be used for the dye solution in PARTS II and III.

PART II DIRECT DYEING

1. Pour 200 ml of the dye solution assigned to you by your teacher into the 400-ml beaker.
2. Put *two* of the samples of cotton cloth from PART I into the dye solution.
3. Bring the cloth and solution to a boil. Boil for three minutes.


You like my tiedyed shirt? Yes sir, that's Action Chemistry for ya!

4. Remove both pieces of cloth from the dye. Use forceps. Rinse the cloth with water. Wring them out. Hang them up to dry.

5. Save the dye solution for PART III.

PART III INDIRECT DYEING

1. Pour 200 ml of aluminum sulfate solution into a second 400-ml beaker.

2. Put the second two pieces of cotton cloth from PART I into the solution. Heat to boiling. Boil for three minutes.

3. Remove the cloths. Use forceps. Rinse the cloths. Wring them out.

4. Pour out the aluminum sulfate solution. Scrub the 400-ml beaker and rinse it well. You will use it in the next step.

5. Place 200 ml of 0.5 *M* ammonium hydroxide solution in the 400-ml beaker. Put the pieces of cloth wrung from the aluminum sulfate solution into this solution. Bring to a boil. Boil for 10 seconds.

6. Remove the cloths. Use forceps. Rinse the cloths with water. Wring out. These two pieces of cloth are now mordanted and ready for dyeing.

7. Place the mordanted samples in the dye solution. Boil for three minutes.

8. Remove the samples with forceps. Rinse with water. Wring out. Hang up to dry.

9. Is there any difference in color in the dry samples from the direct dyeing and those from the indirect dyeing? ______________ If so, how do you account for this? ______________

    ____

10. Attach the dry samples to this page. Label them. Look at them from time to time. Note any fading of the color. If there is any fading, is it faster in the samples from the direct dyeing or in those from the indirect dyeing? ____

    ____

SAMPLES OF CLOTH

Direct dyeing	
Indirect dyeing	

Questions to answer

1. Why is it necessary to remove the sizing before dyeing cloth? ______________________

2. Try to find out whether the particles of a dye are of molecular or colloidal size. ____________

3. The effective mordant is formed in the cloth by the reaction between the aluminum sulfate and the ammonium hydroxide. Write the balanced equation for this reaction.

    ____

    ____

4. What is the name and formula of the mordant?

    ____

Chapter 25

REDUCTION OF AN ORE

Most metals are mined from the earth as oxide, sulfide, or carbonate compounds. The metal compounds as mined, are known as ORES. Metals are obtained from oxide ores by heating the ore with carbon. The process of heating an ore with carbon is called *reduction with carbon.*

In this experiment you will heat lead oxide, PbO, with carbon. You will find out why the process is called reduction with carbon.

Experiment 25.1

Equipment and Supplies per Lab Station

glass tubing, 3-4 inch lengths, fire polished for mouthpieces
blowpipe; burner; spoon or small spatula
rubber tubing to fit blowpipe, 8-12 inch lengths
ring stand with small ring; charcoal blocks; lead monoxide, powdered

What to do

1. Make a small indention about the size of a pea in a charcoal block. Use a spoon handle or a small spatula.

2. Pack lead oxide powder firmly into this cavity.

3. Set the charcoal block on a wire gauze on an iron ring stand.

4. Obtain the following:
   - (a) A blowpipe
   - (b) One 8-12 inch length of rubber tubing
   - (c) One 3-4 inch piece of glass tubing, fire polished at both ends

5. Insert the glass tubing in one end of the rubber tubing; insert the blowpipe in the other end.

6. Light a burner. Adjust the air intake to obtain a blue cone flame.

7. Look at Fig. 118 to see how to hold the blowpipe to direct the flame onto the lead oxide in the charcoal block. Do not inhale while the blowpipe is in your mouth. If you wish to inhale, take the blowpipe out of your mouth.

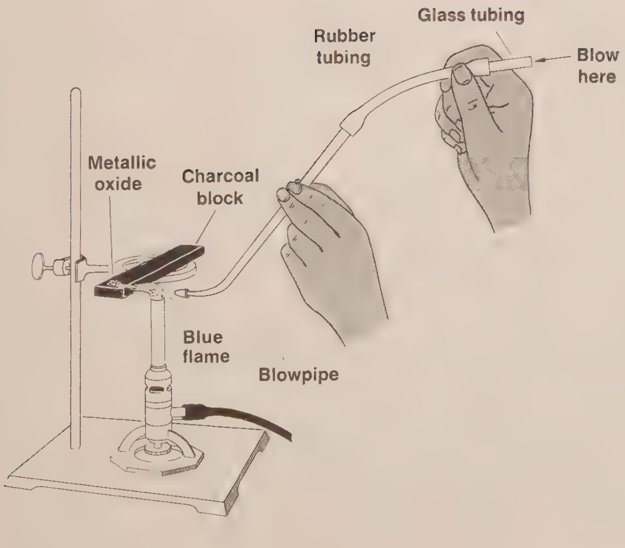
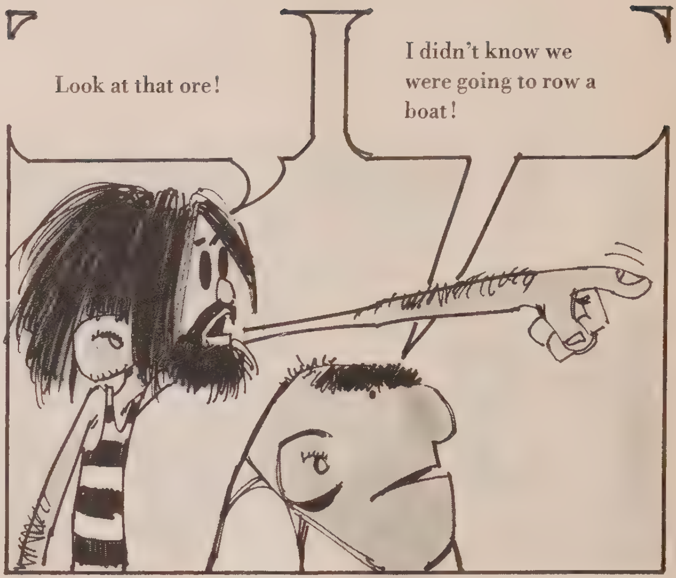


Fig. 118.

8. Heat the lead oxide with the blowpipe for one or two minutes. Write down all of your observations. ______________________________
____

9. Continue to heat the lead oxide until a melted glob is seen. Write down any additional observations as the deposit cools. ____________
____

10. Remove the cool glob of material from the charcoal block. Is it hard or soft? __________
What is its color? ______________ Is it shiny or dull? ______________ Hit the material gently with a hammer. What happens? ________
____
What type of materials have properties similar to those just described? ______________


Look at that ore!
I didn't know we were going to row a boat!

Questions to answer

1. Lead oxide, PbO, reacts with hot carbon, C, (charcoal) to produce a solid and a gas. Complete and balance the following equation.
$PbO + C \rightarrow$ ______________ + ______________
What was the material formed on the charcoal block? ______________________________
2. What is the combining number of lead in lead oxide? ______________ Write the formula for a lead ion. ______________ Does the lead in lead oxide gain or lose electrons to become lead without an electric charge? ______________
3. You learned in Experiment 13.3-A Extension that an atom or ion is reduced when it gains electrons and oxidized when it loses electrons. Is the lead in the lead oxide reduced or oxidized when it changes to uncharged lead? ____
____
Why do you think that the reaction carried out in this experiment is called *reduction with carbon?* ______________________________
____
4. Iron oxide, Fe_2O_3, will react with hot carbon to produce iron. Is the iron in this reaction reduced or oxidized? ______________________________

Roasting a Carbonate Ore

Carbonate and sulfide ores must be changed to oxide ores before they can be reduced with carbon to product metals, as in Experiment 25.1.

The process of changing a metallic carbonate or a metallic sulfide to a metallic oxide is known as ROASTING. The roasting reactions may be written:

metallic carbonate + heat → CO_2 + metallic oxide
metallic sulfide + O_2 + heat → CO_2 + metallic oxide

In this demonstration copper carbonate will be roasted.

Demonstration 25-A

What to do

1. Put one teaspoonful of copper(II) carbonate, $CuCO_3$, in an evaporating dish. Describe the appearance of the copper carbonate. ____________

   ____
2. Set the evaporating dish on an iron ring on a ring stand. Place the apparatus under the hood.
3. Light a burner. Adjust the air intake to obtain a blue cone.
4. Heat the dish and its contents gently at first. Then increase the heat. Heat the carbonate strongly for about five minutes.
5. Cool the dish. Inspect and describe the solid remaining in the dish. ____________

   ____
6. Use a spatula to turn the material over so that you can inspect any substance near the bottom of the dish. Write down any additional observations. ____________

   What do you think the material from the bottom of the dish looks like? ____________
7. Place some of the material from the bottom of the dish on a paper towel. Strike it gently with a hammer. What did you find out from this test? ____________

Equipment and Supplies per Lab Station
evaporating dish; burner; ring stand with ring and wire gauze
copper (II) carbonate, $CuCO_3$, powder
hammer; spatula

Questions to answer

1. Write a balanced equation for the reaction shown in the demonstration. ____________
2. What substance, if discharged directly into the air from the sulfide roasting oven, can cause serious air pollution? ____________

   ____

Extra for experts

See if you can find a way that the air pollutant from sulfide roasting ovens can be removed from the exhaust chimney. You may wish to consult reference books. ____________

____

This is a holdup!
Pay me some dirt,
or else!

Refining Copper — Pay Dirt!

Copper ores often contain zinc and lead-bearing minerals. Some gold or silver may be present in small amounts. After the copper metal has been extracted from the ore, zinc and other metals may still be mixed with the copper. The copper must be purified by removing all other metals. *The process of purifying copper, or any other metal, is known as REFINING.*

In this experiment you will refine a sample of copper metal.

Experiment 25.2

Equipment and Supplies per Lab Station
beaker—250 ml; battery, 6 volt; carbon rod electrode
strip of copper metal electrode; insulated copper wires with alligator clips
electric switch; variable resistance; copper sulfate solution

What to do

1. Obtain about 200 ml of copper sulfate solution. Pour 5 ml of this solution into a test tube. Set the test tube aside for later use in comparing solutions.

2. Set up the apparatus as shown in Fig. 119. Pour the copper sulfate solution into the beaker with the electrodes.
   (a) Which is the positive electrode? ____________
   (b) Which is the negative electrode? ____________
   (c) In which direction will the electric current flow when the switch is closed? ____________
   ____

3. Close the switch. Allow the electricity to flow for two to three minutes.

4. Turn off the electric current by opening the switch.

5. Inspect the electrodes. Describe any change in appearance of
   (a) the copper electrode. ____________
   (b) the carbon electrode. ____________
   (c) What substance has been formed on the negative electrode? ____________
   How do you know this? ____________
   ____

6. Filter some of the copper sulfate solution from the beaker to remove any solid particles that may have dropped into the solution.

7. Compare the color of 5 ml of the filtered solution with the color of the solution in the test tube from step 1. Did the concentration of the copper sulfate solution change during the refining process? ____________
   What makes you think this? ____________
   ____

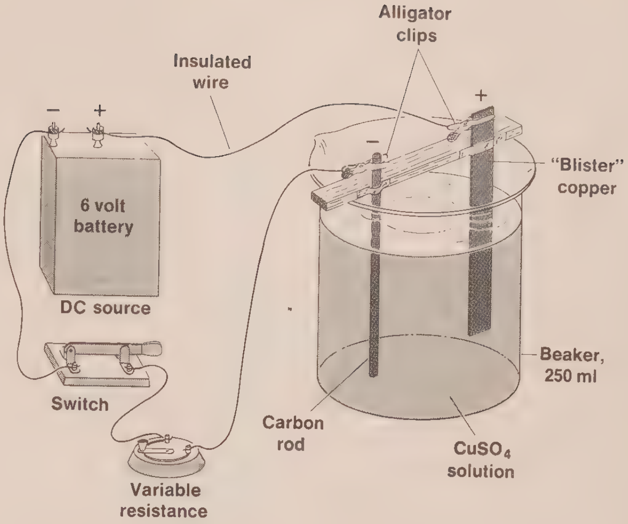


Fig. 119. Set up for the electrolysis of copper.

Questions to answer

1. The battery sends electrons to the negative electrode. What ions in the solution would be attracted to this electrode? ____________
   What occurs at this electrode: oxidation or reduction? (Review Experiment 13.3-A Extension.) ____________
2. What must happen at the positive electrode from which electrons must be sent back to the battery? (Oxidation or reduction?) ____________
   ____
3. Can the sulfate ion exist without its charge? ____
   ____
4. The impure copper extracted from ore is known as *blister copper*. If blister copper were used as the positive electrode, what would happen to the other metals present as impurities? ____________
   ____
5. Why might the sludge at the bottom of an electrolytic refining tank be called "pay dirt?"
   ____
6. In the actual refining process, the negative electrode is originally a very thin sheet of pure copper. Why isn't carbon used? ____________
   ____
7. What happens to the mass of the positive electrode? ____________ The negative electrode? ____________

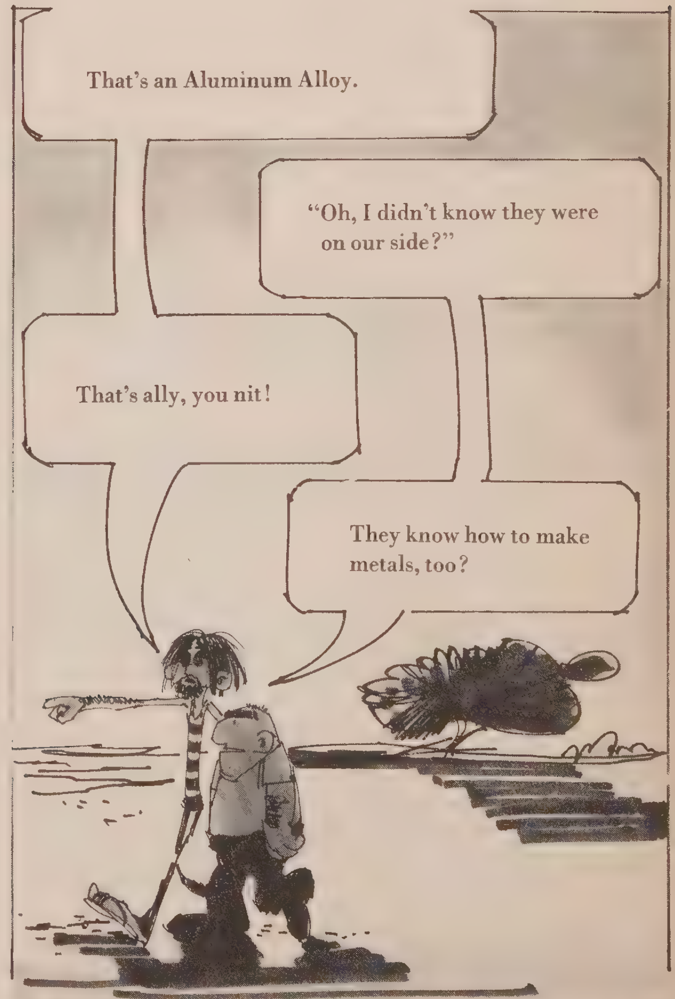
Alloys

The atoms in a pure metal are arranged in a regular crystalline pattern. Layers of atoms rest one layer on top of another so that the atoms in one layer fit into spaces between atoms in the layers above and below.

An alloy is a combination of a metal with one or more other elements. Atoms of the different element have fitted into the crystalline pattern among the atoms of the metal.

An alloy has metallic properties even though it may be a mixture either of two or more different metals or of a metal and a nonmetal.

Why are alloys made? Why not just use the metal? You may find answers to these questions in this experiment.


That's an Aluminum Alloy.
"Oh, I didn't know they were on our side?"
That's ally, you nit!
They know how to make metals, too?

Experiment 25.3

Equipment and Supplies per Lab Station

beaker—250 ml; box, small and shallow; burner; conductivity apparatus

ring stand with ring and wire gauze; tongs

styrofoam balls, three sizes: medium size, 15 or more; large size, 3 to 5; small size, 3 to 5

(2) unknown samples of metal-like substances

What to do

PART I

1. The layered pattern of the metal structure can be represented using small spheres to represent atoms. See Fig. 120. Obtain a shallow box and a number of styrofoam balls all the same size. Completely fill the bottom of the box with the spheres.
2. Place a few incomplete rows of the same kind of sphere in a second layer.
3. Push against the rows in the top layer with the eraser end of a pencil. What happens? ________

   ____
4. Remove the rows in the top layer. Set the spheres aside to use in step 6.
5. Replace one or two spheres in each row of the bottom layer with larger spheres. Place these in an irregular pattern.
6. Replace the top incomplete rows from step 4.
7. Push against the rows in the top layer with the eraser end of a pencil. Record any further observations. ______________

   ____

   How can you account for any new observations?

   ____

   ____
8. Repeat steps 4 and 5. Use smaller spheres instead of larger ones. Replace the top incomplete rows.
9. Push against the rows in the top layer with the eraser end of a pencil. What happens? ________

   Why do you think this happens? ________

   ____

Question to answer

Look up the relative sizes of silver atoms and copper atoms. Sterling silver is an alloy of silver and copper.
Why do you think sterling silver is harder than either copper or silver? ______________

____

____

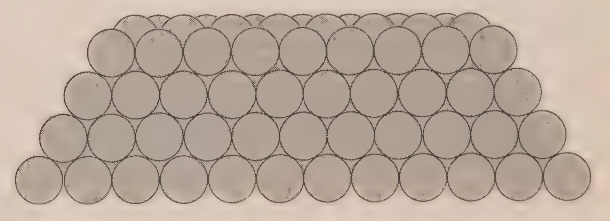


Fig. 120. Atoms in a metal structure represented like layers of cannon balls.

PART II

1. You will be given a sample of material. Inspect the sample. Describe its appearance.

   ____

   ____

2. Use the conductivity apparatus. Is the sample an electrical conductor? ______________
3. Try dissolving the sample in water. Is it soluble? ____ ______________
4. Does the sample have luster? __________

   Is it malleable? ______________
5. Does the sample have the properties of a metal? ______________
6. You will be given a second sample. Describe its appearance. ______________

   ____

7. Check the ability of the sample to conduct electricity. Is it an electrical conductor? _____
8. Put about 100 ml of water in a 250 – ml beaker. Set the beaker on an iron ring stand. Heat the water until it boils.
9. Hold the sample with tongs. Stir the boiling water with the sample as if it were a spoon.

   Write down any observations. __________

   ____

10. Does the sample behave like metal in the hot water? ______________

    ____

    If it is not a metal, what might it be? _____

    ____

    Why do you think so? ______________

    ____

Questions to answer

1. Solder is an alloy of tin and lead. What is solder used for? ______________

   ____

   What property does solder have which is different from that of either tin or lead? _____

   ____

2. Steel contains iron and carbon. Stainless steel also contains chromium. What property does chromium give to the alloy? __________

   ____

APPENDIX

TABLE FROM UNIT 1

DATA TABLE

PROPERTIES OF SOME ELEMENTS

Name	Phase At Room Temperature	Density (g/cm³)	Density (g/l)	Melting Pt. (°C)	Boiling Pt. (°C)	Appearance
Aluminum	solid	2.7		660.2	2467	silvery
Bismuth	solid	9.8		271.3	1560	silvery-white or reddish
Bromine	liquid	3.1		−7.2	58.8	dark red
Chlorine	gas		2.5	−103	−34.6	greenish-yellow
Gold	solid	19.3		1063	2600	yellow
Helium	gas		0.18	−272 (at high pressure)	−268.9	colorless
Hydrogen	gas		0.09	−259	−268.9	colorless
Iodine	solid	4.93		113-114 sublimes	184.4	silvery-blackish solid violet gas
Lead	solid	11.3		327	1515	silvery-black
Magnesium	solid	1.74		651	1107	silvery-white
Oxygen	gas		1.1	−218.4	−183.	colorless
Sulfur	solid	2.0		114.5	444.6	yellow
Zinc	solid	7.14		419.5	907	bluish-white, lusterous

sublimes: turns directly from solid to gas

Helium's melting point is determined under 26 atm of pressure

(Data from *Chemistry and Physics Handbook*, 46th ed.)

TABLE FROM UNITS 2 AND 8

*TABLE OF SOLUBILITIES IN WATER

i -nearly insoluble Ss-slightly soluble s -soluble d -decomposes **n -not isolated	acetate	bromide	carbonate	chloride	hydroxide	iodide	nitrate	oxide	phosphate	sulfate	sulfide
aluminum	s	s	n	s	i	s	s	i	i	s	d
ammonium	s	s	s	s	s	s	s	n	s	s	s
barium	s	s	i	s	s	s	s	s	i	i	d
calcium	s	s	i	s	Ss	s	s	Ss	i	Ss	d
copper(II)	s	s	i	s	i	d	s	i	i	s	i
iron(II)	s	s	i	s	i	s	s	i	i	s	i
iron(III)	s	s	n	s	i	s	s	i	i	Ss	d
lead	s	Ss	i	Ss	i	Ss	s	i	i	i	i
magnesium	s	s	i	s	i	s	s	i	i	s	d
mercury(I)	Ss	i	i	i	n	i	s	i	i	Ss	i
mercury(II)	s	Ss	i	s	i	i	s	i	i	d	i
potassium	s	s	s	s	s	s	s	s	s	s	s
silver	Ss	i	i	i	n	i	s	i	i	Ss	i
sodium	s	s	s	s	s	s	s	d	s	s	s
zinc	s	s	i	s	i	s	s	i	i	s	i

*New York Board of Regents

**"not isolated" means that a compound with the combination indicated has not been found to exist.

TABLE FROM UNIT 3

ACIDS AND THE COMPOUNDS THEY FORM

NAME OF THE ACID	FORMULA OF ACID	MAKES COMPOUNDS CALLED	WHAT GOES INTO COMPOUND
Acetic	$HC_2H_3O_2$	Acetate	$C_2H_3O_2^{-1}$
Carbonic	H_2CO_3	Carbonate	CO_3^{-2}
Chloric	$HClO_3$	Chlorate	ClO_3^{-1}
Hydrobromic	HBr	Bromide	Br^{-1}
Hydrochloric	HCl	Chloride	Cl^{-1}
Hydrofluoric	HF	Fluoride	F^{-1}
Hydroiodic	HI	Iodide	I^{-1}
Hydrosulfuric	H_2S	Sulfide	S^{-2}
Nitric	HNO_3	Nitrate	NO_3^{-1}
Phosphoric	H_3PO_4	Phosphate	PO_4^{-3}
Sulfuric	H_2SO_4	Sulfate	SO_4^{-2}
Sulfurous	H_2SO_3	Sulfite	SO_3^{-2}

THE PERIODIC TABLE

METALS

NON-METALS

Transition elements

I	II											III	IV	V	VI	VII	Noble gases VIII
1 **H** 1.01																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 15.99	9 **F** 18.99	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 51.99	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.71	29 **Cu** 63.55	30 **Zn** 65.37	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** 99	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.4	47 **Ag** 107.87	48 **Cd** 112.40	49 **In** 114.82	50 **Sn** 118.69	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.30
55 **Cs** 132.91	56 **Ba** 137.34	57 **La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.09	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.37	82 **Pb** 207.19	83 **Bi** 208.98	84 **Po** 210	85 **At** 210	86 **Rn** 222
87 **Fr** 223	88 **Ra** 226	89 **Ac** 227															

Rare earth elements

58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** 147	62 **Sm** 150.35	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.92	66 **Dy** 162.50	67 **Ho** 164.930	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
90 **Th** 232.04	91 **Pa** 231	92 **U** 238.03	93 **Np** 237	94 **Pu** 242	95 **Am** 243	96 **Cm** 247	97 **Bk** 249	98 **Cf** 251	99 **Es** 254	100 **Fm** 253	101 **Md** 256	102 **No** 254	103 **Lw** 257

INFORMATION ABOUT COMMONLY-USED ELEMENTS

Name of element	Symbol	Atomic number	Atomic weight	Common oxidation numbers
aluminum	Al	13	26.98	+3
antimony	Sb	51	121.75	+3, +5
argon	Ar	18	39.95	0
arsenic	As	33	74.92	+3, +5
barium	Ba	56	137.34	+2
beryllium	Be	4	9.01	+2
bismuth	Bi	83	208.98	+3
boron	B	5	10.81	+3
bromine	Br	35	79.90	−1
cadmium	Cd	48	112.40	+2
calcium	Ca	20	40.08	+2
carbon	C	6	12.01	+2, +4
cerium	Ce	58	140.12	+3, +4
cesium	Cs	55	132.91	+1
chlorine	Cl	17	35.45	−1, +5
chromium	Cr	24	51.99	+2, +3, +6
cobalt	Co	27	58.93	+2, +3
copper	Cu	29	63.54	+1, +2
fluorine	F	9	18.99	−1
gallium	Ga	31	69.72	+3
germanium	Ge	32	72.59	+4
gold	Au	79	196.97	+1, +3
helium	He	2	4.00	0
hydrogen	H	1	1.01	−1, +1
iodine	I	53	126.90	−1, +5
iron	Fe	26	55.85	+2, +3
krypton	Kr	36	83.80	0
lead	Pb	82	207.19	+2, +4
lithium	Li	3	6.94	+1
magnesium	Mg	12	24.31	+2
manganese	Mn	25	54.94	+2, +4, +7
mercury	Hg	80	200.59	+1, +2
neon	Ne	10	20.18	0
nickel	Ni	28	58.71	+2
nitrogen	N	7	14.01	−3, +3, +5
oxygen	O	8	15.99	−2
phosphorus	P	15	30.97	+3, +5
platinum	Pt	78	195.09	+2, +4
potassium	K	19	39.10	+1
radium	Ra	88	226*	+2
radon	Rn	86	222*	0
scandium	Sc	21	44.96	+3
selenium	Se	34	78.96	−2, +4, +6
silicon	Si	14	28.09	+4
silver	Ag	47	107.87	+1
sodium	Na	11	22.99	+1
strontium	Sr	38	87.62	+2
sulfur	S	16	32.06	±2, +4, +6
tin	Sn	50	118.69	+2, +4
titanium	Ti	22	47.90	+3, +4
tungsten	W	74	183.85	+2, +3, +4, +5, +6
uranium	U	92	238.03	+3, +4, +5, +6
vanadium	V	23	50.94	+2, +3, +4, +5
xenon	Xe	54	131.30	0
zinc	Zn	30	65.37	+2

*Atomic weight of most stable isotopes

TABLE FROM UNIT 7

R_f VALUES FOR SOME FOOD COLORING MATTER

Name	Cellulose Layers R_f value in solvent	
	1	2
yellow 1	0.58	0.51
yellow 2	0.12	0.33
red	0.20	0.55
blue	0.00	0.00
purple	0.19	0.22
scarlet	0.90	0.64